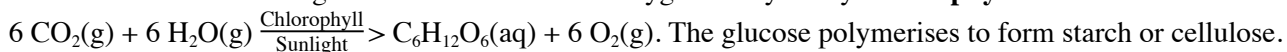


## UNIT 2: ENVIRONMENTAL CHEMISTRY

### Chapter 10: WATER: ESSENTIAL FOR LIFE

Water is essential for life as it:

- provides a transport system for nutrients (digested food), cells & wastes (dissolved carbon dioxide & urea).
- is a good solvent for dissolving foods, wastes and salts and is the only substance that expands when it freezes.
- absorbs  $4.18 \text{ J/}^\circ\text{C/g}$  of heat energy from cells undergoing respiration and perspiration (from the skin).
- provides a cooling system as water evaporates from the skin on animals and from the leaves of trees.
- provides an environment for organisms to live in.
- affects the environment such as tropical rainforests with high rainfalls and the cliffs that are eroded by water.
- is a reactant in \_\_\_\_\_ that occurs in plant leaves during the day to convert Carbon Dioxide in air and Water from the ground into Glucose and Oxygen catalysed by **Chlorophyll** in the leaves.



Water is a covalent polar molecule that can form Hydrogen bonds with 4 other water molecules (1 to each H atom & 1 to each lone pair of  $e^-$  on the \_\_\_\_\_) and to other polar molecules containing Nitrogen, Hydrogen, Oxygen or Fluorine.

The hydrogen bonds are fairly strong Intermolecular Forces and give water a higher than expected BP and MP as more energy is needed to separate the water molecules.

As water freezes, Hydrogen bonds stabilize large air spaces that form between the molecules so water expands and its density decreases. When it melts, the water molecules displace (push out) the air space.

When ice absorbs heat energy, its density increases as it melts since the molecules can get closer and the air spaces \_\_\_\_\_. The heat energy absorbed is called the **Latent Heat**.

- The **Latent Heat of Fusion** is the amount of energy needed to change a fixed amount of solid chemical to liquid. eg. Ice at  $0^\circ\text{C}$  to water. Water's value is higher than other chemicals ( $6 \text{ kJ/mol}$  [ $6 \text{ kJ}$  means  $6000 \text{ Joules}$  of energy and  $\text{mol}$  is the way of measuring the amount of substance]) due to the high energy needed to break the \_\_\_\_\_ between the water molecules in ice to form the liquid.

- The **Latent Heat of Vaporisation** is the amount of energy needed to change a fixed amount of liquid chemical to gas. eg. Water at  $100^\circ\text{C}$  to steam. Water's value is higher than other chemicals ( $44 \text{ kJ/mol}$ ) due to the H-bonds & also explains why water is an effective coolant as heat energy is lost from the skin to the surrounding.

Every substance can store energy. This energy is called the **Specific Heat Capacity** and is the energy needed to increase the temperature of  $1 \text{ gram}$  of the substance by  $1^\circ\text{C}$  using a device called a \_\_\_\_\_.

The value for water is  $4.18 \text{ J g}^{-1} ^\circ\text{C}^{-1}$  which means it needs  $4.18 \text{ J}$  of energy to increase the temperature of  $1 \text{ g}$  of water by  $1^\circ\text{C}$ . The higher the value, the better it can store energy. The value for metals is  $\sim 0.5 \text{ J g}^{-1} ^\circ\text{C}^{-1}$  so metals are better at \_\_\_\_\_ heat than absorbing heat energy.

### Chapter 11: MEASURING SOLUBILITY (like how much salt dissolves in a glass of water)

Water is a good solvent because it is polar so many polar and ionic substances (solutes) will dissolve in it by forming \_\_\_\_\_ which explains why it is difficult to find pure water (distilled water) in nature.

For a substance to dissolve in a suitable liquid (solvent), the bonds (Intermolecular Forces) holding the solute (solid) together must be broken so the particles can separate and form attractions (see below) with the solvent (liquid) to form a solution. eg. A **Solute(s)** that is **Soluble** will dissolve in a **Solvent(l)** and form a **Solution(aq)**.

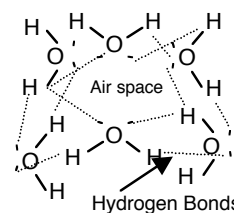
eg. Petrol dissolves in Oil since Dispersion Forces form between both of the non polar hydrocarbon molecules.

eg. Salt dissolves in water since Ion-Dipole interactions form between the salt (ionic) and water (polar or a dipole).

If the bonds holding the solute together are stronger than the bonds that can form with the solvent (as in metals, some ionic solids & giant covalent solids), the substance is \_\_\_\_\_ (will not dissolve in a solvent).

If solutions or solvents are of similar polarities, they can mix or are **Miscible**. eg. Salt solution and water.

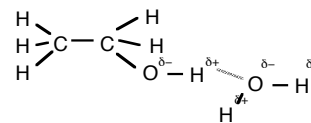
If solutions or solvents are of different polarities, they cannot mix or are **Immiscible**. eg. Water and oil.



The solution is **Homogeneous** if the solute and solvent are so well mixed that they appear the same and cannot be separated and the particles become invisible and the solution becomes clear (transparent).

There are three types of substances that can dissolve in water:

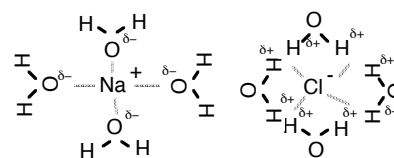
a). Polar molecules (ie. ethanol ( $\text{C}_2\text{H}_5\text{OH}$ )) will dissolve because hydrogen bonds form between the H atom/**O or N atom of a polar group** of a dipole (such as the -OH group of ethanol) and the lone pair of  $e^-$  of an O atom/**H atom** of water. eg. a mixed drink.



b). In the presence of water, some dipoles or polar covalently bonded molecules (there are no ions present) will have the covalent bonds break and form ions in a process called **Ionisation**. Water can do this in small quantities forming a **Hydronium Ion** ( $\text{H}_3\text{O}^+$ ) and a **Hydroxide Ion** ( $\text{OH}^-$ ). eg.  $2 \text{H}_2\text{O}(\text{l}) \rightarrow \text{H}_3\text{O}^+(\text{aq}) + \text{OH}^-(\text{aq})$ .

With substances like HCl, this process occurs readily: eg.  $\text{HCl}(\text{g}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{H}_3\text{O}^+(\text{aq}) + \text{Cl}^-(\text{aq})$  and the ions are said to be **Hydrated** as they are surrounded by water molecules in an aqueous environment.

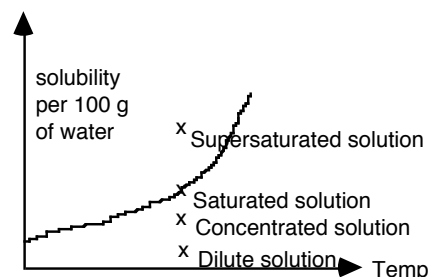
c). Many ionic substances will **Dissociate** (separate) into ions in water and the cations will be attracted to the oxygen atom of water and the anions will be attracted to the hydrogen atom of the water molecules forming **Ion-Dipole Attraction** (ion from the ionic substance and dipole from water being polar). eg. salt in water:  $\text{NaCl}(\text{s}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{Na}^+(\text{aq}) + \text{Cl}^-(\text{aq})$ .



A **Solubility Curve** is used to show the relationship between the solubility of a solute in a solvent and a change in temperature. The solubility of most solid solutes increases as the temperature of the solvent rises.

A **Dilute** solution contains a very small amount of solute dissolved in a large amount of solvent while a **Concentrated** solution contains a larger amount of solute dissolved in the solvent.

A solution in which no more solute can be dissolved at that temperature is a **Saturated** solution. If however, the solution is heated, more solute can be added and if this is then allowed to cool very slowly, the solute will stay dissolved and form a **Supersaturated** solution.



If too much solute is added to a saturated solution, the solute may **Crystallise** or **Precipitate** out of solution as crystals. This also occurs if the solute was dissolved in a hot solvent which is then cooled too quickly.

The **Solubility** of a substance in a solvent is the maximum amount of the solute that can be dissolved in a given quantity of the solvent at a certain temperature. Usually a substance is slightly soluble if the solubility is less than 1 g of solute per 100g of solvent, soluble if greater than 10g/100g and insoluble if less than 0.01g/100g.

Gas solubility **decreases** as the temperature increases but is increased by forcing the gas into solution with pressure. eg.  $\text{CO}_2$  in Soft drinks.

In a solution, the amount of a solute dissolved in a solvent is called **Concentration** (symbol **c**) and has the units:

a). **Percentage by Mass (w/w)** which is the mass (weight, w) in g of solute per \_\_\_\_\_ of solution.

eg. In a 200 g solution containing 28 g of alcohol, the concentration of alcohol as a % w/w is:

$$\text{Concentration, } c = \frac{\text{Mass of Solute [in g]}}{\text{Mass of Solution [in g]}} \times 100 = \frac{28}{200} \times 100 = \mathbf{14 \% \text{ w/w.}}$$

b). **Percentage by Volume (v/v)** which is the volume, v in mL or ( $\text{cm}^3$ ) of solute per 100 mL of solution.

eg. In a 200 mL solution containing 28 mL of alcohol, the alcohol concentration as a % v/v is:

$$\text{Concentration, } c = \frac{\text{Volume of Solute [in mL]}}{\text{Volume of Solution [in mL]}} \times 100 = \frac{28}{200} \times 100 = \mathbf{14 \% \text{ v/v.}}$$

c). **Percentage Mass/Volume (w/v)** which is the mass in g per \_\_\_\_\_ of solution.

eg. In a 200 mL solution containing 28 g of alcohol, the alcohol concentration as a % w/v is:

$$\text{Concentration, } c = \frac{\text{Mass of Solute [in g]}}{\text{Volume of Solution [in mL]}} \times 100 = \frac{28}{200} \times 100 = \mathbf{14 \% \text{ w/v.}}$$

d). **Mass of Solute in g per milliLitre of Solution (g/mL)**. ( $1000 \text{ mL or } 1000 \text{ cm}^3 = 1 \text{ L or } 1 \text{ dm}^3$ ).

eg. In 200 mL solution containing 28 g of alcohol, the alcohol concentration as a mass per milliLitre of solution is:

$$\text{Concentration, } c = \frac{\text{Mass of Solute [in g]}}{\text{Volume of Solution [in mL]}} = \frac{28}{200} = \mathbf{0.14 \text{ g/mL or } \text{g mL}^{-1}}. \quad (\text{To convert L into mL, times by 1000}).$$

e). **Mass of Solute in g per Litre of Solution (g/L).** ( $1000 \text{ mL}$  or  $1000 \text{ cm}^3 = 1 \text{ L}$  or  $1 \text{ dm}^3$ ).

eg. In  $200 \text{ mL}$  solution containing  $28 \text{ g}$  of alcohol, the alcohol concentration as a mass per Litre of solution is:

$$\text{Concentration, } c = \frac{\text{Mass of Solute [in g]}}{\text{Volume of Solution [in L]}} = \frac{28}{0.200} = \mathbf{140 \text{ g/L}} \text{ or } \mathbf{gL^{-1}}. \quad (\text{To convert mL into L, divide by } 1000).$$

f). **Parts per million (ppm)** which is the number of parts of solute in  $1000000$  parts of solution as either the solute's mass in g in \_\_\_\_\_ of solution (or mg in  $1 \text{ kg}$ ) or the solute's volume in mL in  $1000000 \text{ mL}$  of solution.

eg. The concentration of alcohol in ppm for a  $200 \text{ g}$  sample containing  $28 \text{ g}$  of alcohol is:

$$\text{Concentration, } c = \frac{\text{Mass of Solute [in g]}}{\text{Mass of Solution [in g]}} \times 1000,000 = \frac{28}{200} \times 1000,000 = \mathbf{140,000 \text{ ppm}}.$$

g). **Parts per billion (ppb)** which is the number of parts of the substance in  $1000000000$  parts of solution as either the solute's mass in g in  $1 \times 10^9 \text{ g}$  of solution or the solute's volume in mL in  $1 \times 10^9 \text{ mL}$  of solution.

eg. Determine the concentration of alcohol in ppb for a  $200 \text{ g}$  sample containing  $28 \text{ g}$  of alcohol.

$$\text{Concentration, } c = \frac{\text{Mass of Solute [in g]}}{\text{Mass of Solution [in g]}} \times 1000,000,000 = \frac{28}{200} \times 1000,000,000 = \mathbf{140,000,000 \text{ ppb}}.$$

If the question has a measurement in g and another in mL, a density formula (Density,  $d = \frac{\text{Mass}}{\text{Volume}}$  in  $\text{g/mL}$ ) can be used to convert both units of measurement into one unit. If the density is not stated, assume it is  $1.0 \text{ g/mL}$ .

eg. In a  $200 \text{ mL}$  sample of density  $0.92 \text{ g/mL}$  containing  $28 \text{ g}$  alcohol, the concentration of alcohol in ppm is:

Density =  $\frac{\text{mass}}{\text{Volume}}$  so for every  $1.0 \text{ mL}$  there is  $0.92 \text{ g}$  so in  $200 \text{ mL}$  of solution, there is \_\_\_\_\_ g:

$$\text{Concentration} = \frac{\text{Mass of Solute [in g]}}{\text{Volume of Solution [in mL]} \times \text{Density}} \times 1000,000 = \frac{28}{200 \times 0.92} \times 1000,000 = \mathbf{152,174 \text{ ppm}}.$$

h). **MOLARITY (unit M)** measures the amount of solute (in mol) per Litre of solution ( $\text{mol/dm}^3$  or  $\text{mol/L}$ ).

eg. In a  $200 \text{ mL}$  sample containing  $0.028 \text{ mol}$  of alcohol, the concentration of alcohol in Molarity is:

$$\text{Concentration, } c = \frac{\text{Amount of solute [in mol]}}{\text{Volume of Solution [in L]}} = \frac{0.028}{0.20} = \mathbf{0.14 \text{ M}}. \quad (\text{To convert mL into L, divide by } 1000).$$

**To assist converting units, the following flowchart can be used (for masses,  $1 \text{ kg} = 1000 \text{ g}$ ,  $1 \text{ g} = 1000 \text{ mg}$ ):**

**ppb**  $\begin{matrix} <-- \times 1000 \\ \div 1000 --> \end{matrix}$  **ppm**  $\begin{matrix} <-- \times 10000 \\ \div 10000 --> \end{matrix}$  **% w/w or % w/v**  $\begin{matrix} <-- \times 100 \\ \div 100 --> \end{matrix}$  **g/mL**  $\begin{matrix} <-- \div 1000 \\ \times 1000 --> \end{matrix}$  **g/L**  $\begin{matrix} <-- \times M_r \\ \div M_r --> \end{matrix}$  **Molarity (M)**  
 g or mL in  $10^9 \text{ g}$  or mL. g or mL in  $10^6 \text{ g}$  or mL. g in  $100 \text{ g}$  or g in  $100 \text{ mL}$ . g in  $1 \text{ mL}$ . g in  $1000 \text{ mL}$ . mol in  $1 \text{ L}$ .  
 Unless listed, assume the density, ( $d = \frac{m}{V}$ ) is  $1.0 \text{ g/mL}$ , otherwise to convert  $\text{g/mL}$  into  $\text{g/g}$ ,  $\div$  by the density.

eg. Convert  $25 \text{ ppm NaCl}$  into \_\_\_\_\_  $\text{ppb}$  into \_\_\_\_\_  $\%$  w/w into \_\_\_\_\_  $\text{g/mL}$ .

Working Space:

The above formulae can be rearranged to find the amount of solute in a solution with a certain concentration.

eg. Find the mass of Potassium expected in a  $100 \text{ mL}$  solution containing  $14 \text{ gL}^{-1}$  of Potassium.

$$\text{Mass of Solute (in g)} = \text{Concentration (in } \text{gL}^{-1}) \times \text{Volume (in L)} \quad (\text{To convert mL into L, divide by } 1000).$$

$$\text{Mass of Potassium} = 14 \times 100 \div 1000 = \mathbf{1.4 \text{ g}}.$$

Be Careful: If asked for the mass of  $\text{H}_2\text{O}$  to be added, subtract the mass of solute from the mass of solution.

eg. For a  $9.0 \%$  solution, then  $9.0 \text{ g}$  of solute must be added to  $91 \text{ g}$  of  $\text{H}_2\text{O}$  to form  $100 \text{ g}$  of solution.

**Calculating Amounts (in mol) in Solutions:** Always use the mol formula – be careful using the correct units.

| For solutes in Solutions  | For solutes in Solutions  |
|---|---|
| a). Determine the <b>amount (in mol)</b> of NaCl present in $500 \text{ mL}$ of $0.25 \text{ M}$ NaCl solution.<br><br>$V = 500 \text{ mL} = 0.500 \text{ L}$ ( $\text{mL} \div 1000 = \text{L}$ ), $c = 0.25 \text{ M}$ , Find $n$<br>$n(\text{NaCl}) = \text{Conc. } c \text{ in M} \times \text{Volume, } V \text{ in L}$<br>$=$ _____<br>$n(\text{NaCl}) =$ _____ | b). Determine the <b>concentration of NaCl</b> present in a $2.0 \text{ L}$ solution containing $0.12 \text{ mol}$ of NaCl.<br><br>$n = 0.12, V = 2.0 \text{ L}$ , Find $c$<br>$[\text{NaCl}] = \frac{n(\text{NaCl}) \text{ in mol}}{V \text{ in L}}$ The [ ] represents the conc.<br>$= \frac{0.12}{2.0}$ so the $[\text{NaCl}] = 0.060 \text{ M}$ . |

Some calculations involve the combination of 2 or more Mol formula to calculate a measurement. Use the measurements stated to calculate the amount in mol ( $n$ ) and then use this to find the required measurement. If part of a solution is removed, the amount in mol will be decreased but the concentration does not change.

| Determining another unit using the Amount   | Determining another unit using the Amount in mol  |
|---|---|
| <p>Determine the <b>mass of Glucose</b>, <math>C_6H_{12}O_6</math> present in 2.0 L of 0.25 M Glucose solution</p> <p><math>V = 2.0 \text{ L}, c = 0.25 \text{ M}</math>, Find <math>n</math> and then find <math>m</math></p> $n(C_6H_{12}O_6) = c \times V$ $= 0.25 \times 2.0$ $= 0.50 \text{ mol}$ $m(C_6H_{12}O_6) = n(C_6H_{12}O_6) \times M_r$ $= 0.50 \times 180$ $m(C_6H_{12}O_6) = 90 \text{ g.}$ | <p>Determine the <b>number of Glucose</b>, <math>C_6H_{12}O_6</math> <b>molecules</b> present in 2.0 L of 0.50 M Glucose solution</p> <p><math>V = 2.0 \text{ L}, c = 0.25 \text{ M}</math>, Find <math>n</math> and then find number of molecules</p> $n(C_6H_{12}O_6) = c \times V$ $= 0.50 \times 2.0$ $= 1.0 \text{ mol}$ $\text{No. of } C_6H_{12}O_6 = n(C_6H_{12}O_6) \times N_a (6.02 \times 10^{23})$ $= 1 \times 6.02 \times 10^{23}$ $= 6.0 \times 10^{23} \text{ } C_6H_{12}O_6 \text{ molecules.}$ |

When a solution is **Diluted**, more of the \_\_\_\_\_ (eg. water) is added to the solution so the amount in mol is constant but the concentration decreases as the volume has increased ( $c$  decreases =  $\frac{n \text{ is constant}}{V \text{ increases}}$ ).

| Diluting a Solution   | Mixing Different Solutions   |
|---|--|
| <p>There are two different questions that can be provided:</p> <ol style="list-style-type: none"> <li>1. The solution is diluted to a final volume.</li> <li>2. A certain volume of solvent is added to the solution.</li> </ol> <p>Use the formula: <math>c_1 \times V_1 = c_2 \times V_2</math> where <math>V_2</math> is the Final (total) volume. Ensure Volumes are in the same unit.</p> <p>eg. Find the new <b>NaCl concentration</b> when 800 mL of water is added to 200 mL of 0.50 M NaCl solution.</p> <p><math>V_1 = 200 \text{ mL}, c_1 = 0.50 \text{ M}, V_2 = 800 + 200 = 1000 \text{ mL}</math>. Find <math>c_2</math></p> $n(\text{NaCl}) \Rightarrow c_1 \times V_1 = c_2 \times V_2$ $\Rightarrow 0.50 \times 200 (\div 1000) = c_2 \times 1000 (\div 1000)$ $C_2 = \frac{0.50 \times 200}{1000} = 0.10 \text{ M.}$ <p>The division by 1000 does not need to be done as it can be cancelled on both sides.</p> | <p>Use the formula: <math>c_1 \times V_1 + c_2 \times V_2 = c_3 \times V_3</math></p> <p>Ensure the Volumes are in the same unit.</p> <p>Since the volumes and concentration are in the same unit and being compared <u>without</u> calculating the amount in mol, they do not need to be in L.</p> <p>eg. Find the new <b>NaCl concentration</b> when 200 mL of 0.80 M NaCl solution is added to 300 mL of 0.60 M NaCl solution.</p> <p><math>V_1 = 200 \text{ mL}, c_1 = 0.80 \text{ M}, V_2 = 300 \text{ mL}, c_2 = 0.60 \text{ M}</math>,<br/> <math>V_3 = 200 + 300 = 500 \text{ mL}</math> Find <math>c_3</math></p> $n(\text{NaCl}) \Rightarrow c_1 \times V_1 + c_2 \times V_2 = c_3 \times V_3$ $\Rightarrow 0.80 \times 200 + 0.60 \times 300 = c_3 \times 500$ $\Rightarrow 160 + 180 = 340 = C_3 \times 500 \text{ so } C_3 = \frac{340}{500}$ $C_3 = 0.68 \text{ M.}$ |

### Reactions In Solutions: Balancing Equations

A balanced chemical equation lists the chemical formula and states of the reactants and products in the reaction.

eg. Lead Nitrate reacts with Aluminium Chloride.

**Step 1: Use a table of ions to determine the symbols and charges of each ion (Remember the ions in Unit 1).**

**Write the metal cation (+ ions) before the non-metal anion(s) (- ions) except for acetate ( $CH_3COO^-$ ).**

eg. Lead ion = Nitrate ion = Aluminium ion = Chloride ion =

**Step 2: Write down ONE of each ion in the reactants and balance the charges of the ions so they are equal by writing the required coefficient (you can cross multiply the charges) after the ion to the BOTTOM RIGHT.**

**For a group of atoms in ions (more than 1 capital letter), place the group in a bracket. eg.  $W(OH)_2$   $(NH_4)_2O$ .**

eg.

**Step 3: When reacting the cations and anions swap partners so write one of each of the ions in the products.**

eg.

**Step 4: Repeat Step 2 to balance the charges of the products and write the chemical formula of the products.**

**NOTE:  $H^+$  and  $OH^-$  or  $O^{2-}$  react to form  $H_2O$ ;  $H^+$  and  $CO_3^{2-}$  react to form  $H_2O$  and  $CO_2$  not  $H_2CO_3$ .**

eg.

**Step 5: Balance the number of atoms on each side of --> (Reactants (LHS) --> Products (RHS)) by writing the required coefficient in FRONT of the chemical (the formula cannot be altered). Include all states of species.**

eg.

A **Precipitation** reaction occurs when two clear solutions are mixed and an insoluble compound (solid) called a **Precipitate** is produced. The precipitate can be isolated by filtering the solution and collecting the **Residue** (s) in the filter paper while the **Filtrate** solution will contain soluble (aq) \_\_\_\_\_ **Ions** that do not react.

A **Solubility Table** is used to determine the solubility of ionic substances. Since most hydrocarbons are non polar, they are insoluble in water (except acetate;  $\text{CH}_3\text{COO}^-(\text{aq})$  and alcohols;  $\text{C}_n\text{H}_{2n+1}\text{OH}$ , which both contain O).

*All anions combined with  $\text{H}^+$ ,  $\text{Li}^+$ ,  $\text{Na}^+$ ,  $\text{K}^+$ ,  $\text{Rb}^+$ ,  $\text{Cs}^+$  (Group 1 metals),  $\text{NH}_4^+$  are SOLUBLE (aq) AND NOT (s).*

*All cations combined with  $\text{NO}_3^-$ ,  $\text{HCO}_3^-$  &  $\text{CH}_3\text{COO}^-$  (acetate) are SOLUBLE (aq).*

*$\text{Cl}^-$ ,  $\text{Br}^-$ ,  $\text{I}^-$  substances are SOLUBLE with all cations except  $\text{Ag}^+$ ,  $\text{Hg}_2^{2+}$  &  $\text{Pb}^{2+}$  which are INSOLUBLE (s).*

*$\text{SO}_4^{2-}$  substances are SOLUBLE with all cations except  $\text{Ca}^{2+}$ ,  $\text{Sr}^{2+}$ ,  $\text{Ba}^{2+}$  &  $\text{Pb}^{2+}$  which are INSOLUBLE (s).*

*$\text{OH}^-$  substances are INSOLUBLE (s) except for  $\text{H}^+$ ,  $\text{Li}^+$ ,  $\text{Na}^+$ ,  $\text{K}^+$ ,  $\text{Rb}^+$ ,  $\text{Cs}^+$ ,  $\text{NH}_4^+$ ,  $\text{Sr}^{2+}$  &  $\text{Ba}^{2+}$  which are SOLUBLE (aq).*

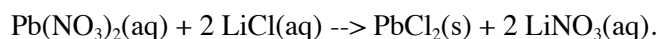
If one of the ions is soluble, the **WHOLE CHEMICAL** is soluble (aq). eg.  $\text{Li}_2\text{CO}_3$  is (aq) as  $\text{Li}^+$  is always (aq)

eg. Which of the following are insoluble:

- a). Calcium Carbonate.      b). Potassium Hydroxide.      c). Silver Chloride.      d). Iron Nitrate.

### **Writing Ionic Equations**      (only include the species that change state: (s) – (l) – (g) – (aq))

eg. Solutions of Lead Nitrate & Lithium Chloride are placed in a beaker and after reacting the products are filtered. The Lithium Nitrate solution is the **Filtrate** and the solid **Residue** precipitate in the filter paper is Lead Chloride:



### **Step 1: Use the Molecular equation to identify the species that have changed state: (aq) <-> (s) <-> (l) <-> (g)**

Use the Solubility Table to identify precipitates (s) as well as other common (covalent) chemicals such as:

Note: If the chemical is (s), (l) or (g), it cannot be split. eg. use  $\text{CaCO}_3(\text{s})$  not  $\text{Ca}^{2+}(\text{s})$ ,  $\text{CO}_3^{2-}(\text{s})$ , use  $\text{H}_2(\text{g})$  not  $2 \text{H}^+(\text{g})$ .

$\text{CO}_2(\text{g})$  and  $\text{H}_2\text{O}(\text{l})$  (from  $\text{H}^+(\text{aq})$  and  $\text{CO}_3^{2-}(\text{aq})$ ),       $\text{H}_2\text{S}(\text{g})$  (from  $\text{H}^+(\text{aq})$  and  $\text{S}^{2-}(\text{aq})$ ),

$\text{H}_2\text{O}(\text{l})$  (from  $\text{H}^+(\text{aq})$  and  $\text{OH}^-(\text{aq})$  or  $\text{H}^+(\text{aq})$  and  $\text{O}^{2-}(\text{aq})$ ),  $\text{H}_2(\text{g})$  from  $2 \text{H}^+(\text{aq})$  reacting with some metals.

eg.  $\text{Pb}(\text{NO}_3)_2(\text{aq}) + 2 \text{LiCl}(\text{aq}) \rightarrow \text{PbCl}_2(\text{s}) + 2 \text{LiNO}_3(\text{aq})$ . If (aq), the species are ionic ( $\text{Pb}^{2+}$  and  $\text{Cl}^-$ )

### **Step 2: Write the species that have changed state with their states (and their charges if there are (aq) ions) and balance the equation by writing the required coefficient in front of the species.**

Note: The species that do not change state are called **Spectator** ions and are not included in the ionic equation.

eg.  $\text{Pb}^{2+}(\text{aq}) + \text{Cl}^-(\text{aq}) \rightarrow \text{PbCl}_2(\text{s})$  but check the equation is balanced so it becomes:  $\text{Pb}^{2+}(\text{aq}) + 2 \text{Cl}^-(\text{aq}) \rightarrow \text{PbCl}_2(\text{s})$

eg. Solid Calcium Carbonate is placed in a solution of Hydrochloric acid according to the equation:



eg:

### **TREATMENT OF DRINKING WATER.**      (like making yukky water yummy)

Waste Water (domestic sewage and industrial waste water) undergoes Primary Treatment of passing the sewage through a screen to remove large objects, then allowing the sludge to settle and skimming the scum off the surface. Drinking water is treated in Australia to remove bacteria, suspended solids, color and odour by:

- Flocculation** which involves adding a chemical such as Alum ( $\text{Al}_2(\text{SO}_4)_3$ ) to make smaller particles clump together to form a gelatinous precipitate called a **Floc** such as Aluminium Hydroxide  $\text{Al}(\text{OH})_3$  which settles to the bottom. If Alum does not work, Lime ( $\text{Ca}(\text{OH})_2$ ) is added to provide a source of Hydroxide ions ( $\text{OH}^-$ ). eg.  $\text{Ca}(\text{OH})_2(\text{s})$  in water  $\rightarrow \text{Ca}^{2+}(\text{aq}) + 2 \text{OH}^-(\text{aq})$  which then reacts with Alum,  $\text{Al}^{3+}(\text{aq}) + 3 \text{OH}^-(\text{aq}) \rightarrow \text{Al}(\text{OH})_3(\text{s})$ . A solid in the solution forms a **Colloid** (a substance suspended in another) that eventually settles to the bottom.
- Settling** allows the water to stand so that the floc (solid particles) settle to the bottom forming a sludge.
- Filtering** the water through a bed of sand over gravel to remove any remaining suspended material.
- Chlorinating** the water with Chlorine gas which provides the Hypochlorite ion ( $\text{OCl}^-$ ) ion to destroy bacteria.
- Fluoridating** the water with Sodium Fluoride or Calcium Fluoride to reduces the incidence of tooth decay.

Sometimes the water is found to be hard which is caused by presence of metal ions such as Calcium, Iron or Magnesium reacting with a substance in soaps called Sodium Stearate ( $C_{17}H_{35}COO^-Na^+$ ) to produce a **Scum**.

There are two types of water hardness:

a). Temporary hardness due to the presence of Calcium and Magnesium which react with the Carbonate ion produced when soluble Hydrogen Carbonate (Bicarbonate) ions present decompose:



b). Permanent hardness due to the presence of soluble Magnesium and Calcium Sulfates and Chlorides in the presence of soap ions which forms an insoluble scum:  $2 C_{17}H_{35}COO^-(aq) + Ca^{2+}(aq) \rightarrow Ca(C_{17}H_{35}COO)_2(s)$ .

Hard Water can be treated by:

a). **Distillation** which involves boiling it and collecting the pure water vapour as it condenses in a condenser into a liquid. This is the way \_\_\_\_\_ water is produced.

b). **Reverse Osmosis** which involves using pressure to force pure water molecules through a membrane away from a salt solution (and not towards it) but the pressure must be controlled to avoid breaking the membrane.

c). **Ion Exchange** which involves passing the water down tubes containing ionic polymers that attract the cations & exchange them with  $H^+$  ions or attract the anions & exchange them with  $OH^-$  ions that don't cause the hardness.

d). **Desalination** involves removing the salt from salt water to produce pure water.

### Chapter 13: INTRODUCING ACIDS AND BASES (like vinegar and bicarb)

**Acids** are usually sour tasting, can change the colour of Indicators (dyes) and neutralize bases. Examples include Hydrochloric Acid,  $HCl$  - for cleaning bricks, Sulfuric Acid,  $H_2SO_4$  - battery acid, Nitric Acid,  $HNO_3$  - dissolves Gold & Silver, Phosphoric Acid,  $H_3PO_4$  - cleans jewellery and Ethanoic or Acetic Acid,  $CH_3COOH$  - vinegar.

**Bases** or **Alkalis** also change the colour of indicators and neutralize acids. Examples include Ammonia,  $NH_3$  - cleaners, Calcium Hydroxide,  $Ca(OH)_2$  - used in plaster and cement, Sodium Hydroxide,  $NaOH$  - drain/oven cleaners and Magnesium Hydroxide,  $Mg(OH)_2$  - used in aiding indigestion.

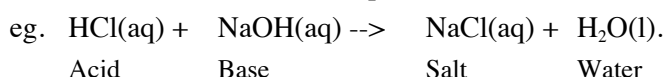
eg. **Litmus** and **Universal Indicator** (a mixture of indicators) turn \_\_\_\_\_ in acid and purple in a base.

Universal Indicator can produce different colours depending upon the concentration of the acid or base.

**Reactions of Acids & Bases:** Hint: Look for products:  $H_2O$  or/ &  $CO_2$ ,  $H_2$  or  $H_2S$ , then determine other product

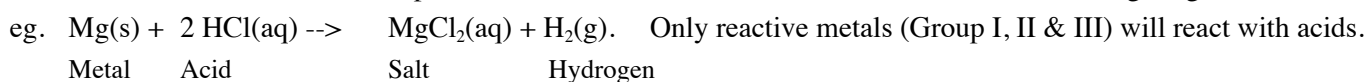
a). **Neutralisation Of Acids Or Bases:** -

Acids will react with Bases to produce a neutral solution of an Ionic Salt and \_\_\_\_\_.



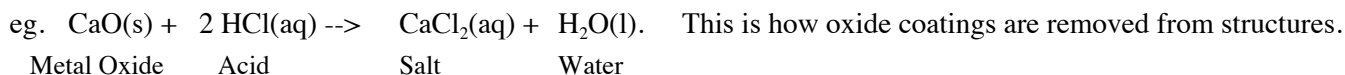
b). **Reaction of Acids with Metals:**

Acids react with certain Metals to produce a neutral solution of an Ionic Salt as well as releasing  $H_2$  gas.



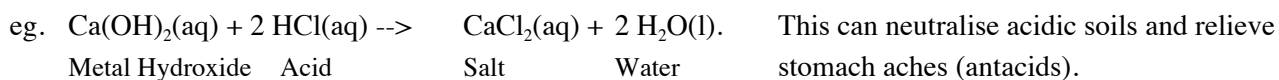
c). **Reaction of Acids with Metal Oxides:**

Acids react with certain Metal Oxides to produce a neutral solution of an Ionic Salt and \_\_\_\_\_.



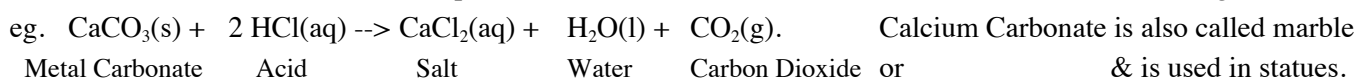
d). **Reaction of Acids with Metal Hydroxides:**

Acids react with certain Metal Hydroxides to produce a neutral solution of an Ionic Salt and  $H_2O$ .



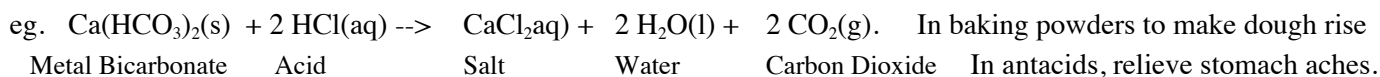
e). **Reaction of Acids with Metal Carbonates:**

Acids react with Metal Carbonates to produce a neutral solution of an Ionic Salt,  $H_2O$  and release  $CO_2$  gas.

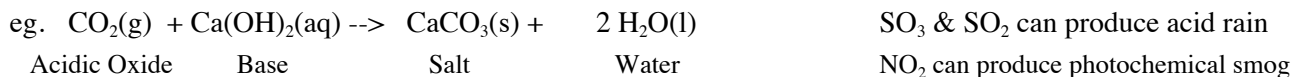


**f). Reaction of Acids with Metal Bicarbonates (Hydrogen Carbonates):**

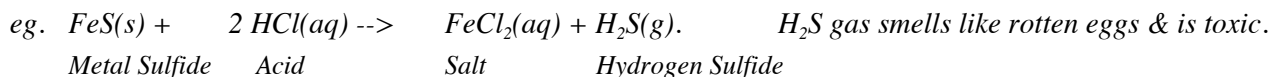
Acids react with Metal Bicarbonates to produce a neutral solution of an Ionic Salt & H<sub>2</sub>O and release CO<sub>2</sub> gas.

**g). Reaction of an Acidic Oxide with a Base:**

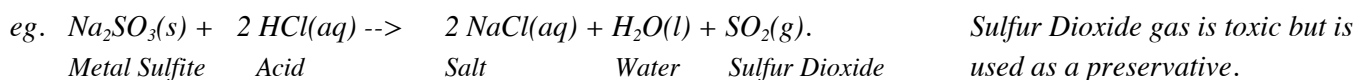
Bases react with Acidic oxides (SO<sub>3</sub>, SO<sub>2</sub>, CO<sub>2</sub>, P<sub>4</sub>O<sub>10</sub>, NO<sub>2</sub>) to produce an Ionic salt and water.

**h). Reaction of Acids with Metal Sulfides:**

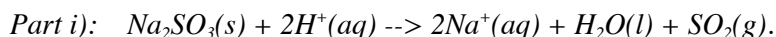
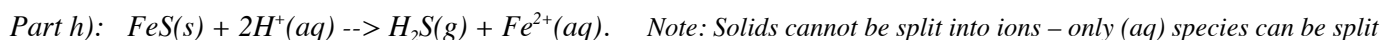
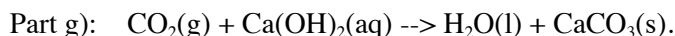
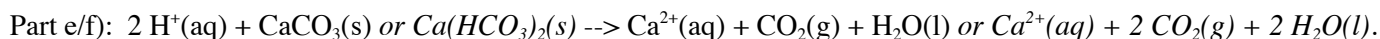
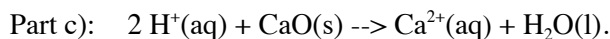
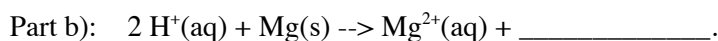
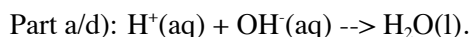
Acids react with some Metal Sulfides to produce a solution of an \_\_\_\_\_ & release H<sub>2</sub>S gas.

**i). Reaction of Acids with Metal Sulfites:**

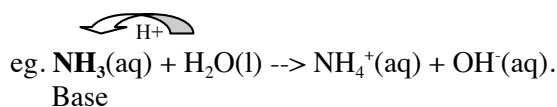
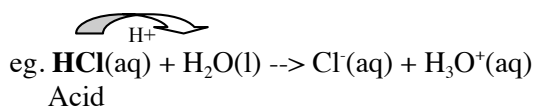
Acids react with some Metal Sulfites to produce a neutral solution of an Ionic Salt & H<sub>2</sub>O and release SO<sub>2</sub> gas.



When writing Ionic Equations, include species that change state, don't split (s), (l) & (g) but ignore Spectator Ions:

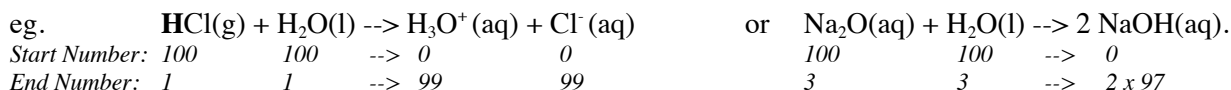
**Chapter 14: ACIDS AND BASES**

**Lowry And Bronsted** defined an **Acid** as a substance able to donate protons (H<sup>+</sup>) and a **Base** or **Alkaline** as a substance that can \_\_\_\_\_ protons (H<sup>+</sup>). To show an acid/base donating/accepting a proton in a reaction where no other reactant is mentioned, water, H<sub>2</sub>O(l) can be used. Common acids start with H and bases end in OH.



The **H<sup>+</sup>** ion can be bound to water to form the **Hydronium Ion (H<sub>3</sub>O<sup>+</sup>)** so **H<sup>+</sup>** is chemically similar to **H<sub>3</sub>O<sup>+</sup>**.

An acid/base that ionises easily to lose/gain a proton easily is a **Strong Acid/Base**. A normal arrow (→) is used to show the reaction goes to completion. Strong acids start with H and bases may end in OH. eg. HCl, Hydrochloric acid) or a base (Sodium Oxide, Na<sub>2</sub>O) are strong since they can fully donate or accept their protons, (H<sup>+</sup>).



An acid/base that cannot ionise easily or lose/gain a proton easily is a **Weak Acid/Base**. A \_\_\_\_\_ arrow is used to show the reaction is incomplete and almost all of the original molecules will remain in solution. eg. Vinegar or Acetic acid, CH<sub>3</sub>COOH or the base Ammonia, NH<sub>3</sub> since they incompletely donate or accept their H<sup>+</sup>.



Strongest Acid: HCl, H<sub>2</sub>SO<sub>4</sub>, HNO<sub>3</sub>, H<sub>3</sub>O<sup>+</sup>, HSO<sub>4</sub><sup>-</sup>, H<sub>3</sub>PO<sub>4</sub>, HF, CH<sub>3</sub>COOH, H<sub>2</sub>CO<sub>3</sub>, H<sub>2</sub>S, H<sub>2</sub>PO<sub>4</sub><sup>-</sup>, H<sub>2</sub>O Weakest.

Weakest Base: Cl<sup>-</sup>, HSO<sub>4</sub><sup>-</sup>, NO<sub>3</sub><sup>-</sup>, H<sub>2</sub>O, SO<sub>4</sub><sup>2-</sup>, H<sub>2</sub>PO<sub>4</sub><sup>-</sup>, F<sup>-</sup>, CH<sub>3</sub>COO<sup>-</sup>, HCO<sub>3</sub><sup>-</sup>, HS<sup>-</sup>, HPO<sub>4</sub><sup>2-</sup>, OH<sup>-</sup> Strongest.

The strength of an acid or base is not related to its concentration (concentrated or dilute).

eg. A small amount of HCl will completely ionise in water to form a dilute strong acid but acetic acid is a polar molecule that does not totally **ionise** in water so a concentrated solution will form a concentrated weak acid solution.

An acid that can donate only one proton is a **Monoprotic** acid. eg.  $\text{HCl(aq)} + \text{H}_2\text{O(l)} \rightarrow \text{Cl}^-\text{(aq)} + \text{H}_3\text{O}^+\text{(aq)}$ .

An acid that can donate 2 protons is a \_\_\_\_\_ acid. The 1<sup>st</sup>  $\text{H}^+$  donation to  $\text{H}_2\text{O}$  is usually complete: eg.  $\text{H}_2\text{SO}_4\text{(aq)} + \text{H}_2\text{O(l)} \rightarrow \text{HSO}_4^-\text{(aq)} + \text{H}_3\text{O}^+\text{(aq)}$  but the 2nd  $\text{H}^+$  being donated by  $\text{HSO}_4^-\text{(aq)}$  to another  $\text{H}_2\text{O}$  molecule is an incomplete reaction:  $\text{HSO}_4^-\text{(aq)} + \text{H}_2\text{O(l)} \rightleftharpoons \text{SO}_4^{2-}\text{(aq)} + \text{H}_3\text{O}^+\text{(aq)}$ .

An acid that can donate 3 protons is a **Triprotic** acid. The 1<sup>st</sup> donation to  $\text{H}_2\text{O}$  is usually \_\_\_\_\_:

eg.  $\text{H}_3\text{PO}_4\text{(aq)} + \text{H}_2\text{O(l)} \rightarrow \text{H}_2\text{PO}_4^-\text{(aq)} + \text{H}_3\text{O}^+\text{(aq)}$  but the next 2  $\text{H}^+$  donations stages are both incomplete:

The  $\text{H}_2\text{PO}_4^-\text{(aq)}$  can donate a  $\text{H}^+$  to another  $\text{H}_2\text{O}$  molecule;  $\text{H}_2\text{PO}_4^-\text{(aq)} + \text{H}_2\text{O(l)} \rightleftharpoons \text{HPO}_4^{2-}\text{(aq)} + \text{H}_3\text{O}^+\text{(aq)}$ .

The  $\text{HPO}_4^{2-}\text{(aq)}$  can donate the last  $\text{H}^+$  to another  $\text{H}_2\text{O}$  molecule;  $\text{HPO}_4^{2-}\text{(aq)} + \text{H}_2\text{O(l)} \rightleftharpoons \text{PO}_4^{3-}\text{(aq)} + \text{H}_3\text{O}^+\text{(aq)}$ .

**Amphiprotic** or **Amphoteric** species or **Ampholytes** can donate AND accept protons,  $\text{H}^+$ . eg.  $\text{HSO}_4^-\text{(aq)}$ ,  $\text{H}_2\text{PO}_4^-$ .  
eg.  $\text{HSO}_4^-\text{(aq)} + \text{H}_2\text{O(l)} \rightleftharpoons \text{SO}_4^{2-}\text{(aq)} + \text{H}_3\text{O}^+\text{(aq)}$  AND  $\text{HSO}_4^-\text{(aq)} + \text{H}_2\text{O(l)} \rightleftharpoons \text{H}_2\text{SO}_4\text{(aq)} + \text{OH}^-\text{(aq)}$ .

If two related chemicals differ by only 1  $\text{H}^+$  they form a **Conjugate Acid/Base Pair** where the chemical with the extra  $\text{H}^+$  is the acid and the other chemical is the base. eg.  $\text{H}_2\text{SO}_4\text{(aq)}$ /  $\text{HSO}_4^-\text{(aq)}$  but not  $\text{H}_2\text{SO}_4\text{(aq)}/\text{SO}_4^{2-}\text{(aq)}$ .

In a Conjugate Acid/Base pair, if the acid/base is strong, then the conjugate base/acid is weak and vice versa.

Water also forms Conjugate Acid/Base pairs but only a 1 in every  $10^7$  molecules self ionize to donate or accept a  $\text{H}^+$  (1 in  $10^7$ ). eg.  $\text{H}_3\text{O}^+/\text{H}_2\text{O}$  and  $\text{H}_2\text{O}/\text{OH}^-$  are shown in the equation:  $2 \text{H}_2\text{O(l)} \rightleftharpoons \text{Hydronium ion, H}_3\text{O}^+\text{(aq)} + \text{OH}^-\text{(aq)}$ .

In neutral pure water at 25 °C, the concentration of the  $\text{OH}^-$  and  $\text{H}_3\text{O}^+$  ions are equal at  $1 \times 10^{-7} \text{ M}$ .

When these values are multiplied together, the result for any solution at 25 °C is always \_\_\_\_\_  $\text{M}^2$ .

eg.  $[\text{H}_3\text{O}^+ \text{ or } \text{H}^+] \times [\text{OH}^-] = 1.0 \times 10^{-14} \text{ M}^2$  where [ ] measures the concentration in mol/L or Molarity, M.

**Acidic** solutions have a greater concentration of  $\text{H}_3\text{O}^+$  than  $\text{OH}^-$  so  $[\text{H}_3\text{O}^+] > 1 \times 10^{-7} \text{ M}$  and  $[\text{OH}^-] < 1 \times 10^{-7} \text{ M}$ .

**Alkaline** solutions have a lower concentration of  $\text{H}_3\text{O}^+$  than  $\text{OH}^-$  so  $[\text{H}_3\text{O}^+] < 1 \times 10^{-7} \text{ M}$  and  $[\text{OH}^-] > 1 \times 10^{-7} \text{ M}$ .

A scale called the **pH** scale (power of hydrogen) is used to determine the acidity of a solution and has no unit.

**pH = - Log $[\text{H}_3\text{O}^+]$** . This law can be rearranged so that  $[\text{H}^+]$  or  $[\text{H}_3\text{O}^+] = 1 \times 10^{-\text{pH}} \text{ M}$ . **Hint: Use LOG not Ln**

Neutral solutions have a pH = 7, acidic solutions have a pH \_\_\_\_\_ 7 and basic solutions have a pH > 7.

**pH scale:** (acidic) 0 --> less acidic (diluted or weak) 7 (neutral) <-- less alkaline (diluted or weak) 14 (alkaline)

**Universal Indicator:** Pink Orange Yellow Green Aqua Blue Purple

eg. Find the **pH** of 0.010 M  $\text{H}_2\text{SO}_4$ .

In 1  $\text{H}_2\text{SO}_4 =$  \_\_\_\_\_  $\text{H}^+$  so the  $[\text{H}^+ \text{ or } \text{H}_3\text{O}^+] =$  \_\_\_\_\_ M.

pH = \_\_\_\_\_ **Hint: Use pH = - Log $_{10}$   $[\text{H}_3\text{O}^+]$ .**

eg. Find the **pH** of 0.010 M NaOH.

In 1 NaOH there is 1  $\text{OH}^-$  so the  $[\text{OH}^-] = 0.010 \text{ M}$  but there is no  $\text{H}^+$ .

Use  $[\text{OH}^-] \times [\text{H}^+] = 1 \times 10^{-14} \text{ M}^2$  to find  $[\text{H}^+]$ , so  $[\text{H}^+] = \frac{1 \times 10^{-14}}{0.01} = 1.0 \times 10^{-12} \text{ M}$

pH = - log $_{10}$   $[1.0 \times 10^{-12}] = 12$  Since the chemical was a base, the pH is > 7.

eg. Find the **concentration** of  $\text{H}^+$  in a solution of pH = 3.0.

$[\text{H}^+] =$  \_\_\_\_\_ **Hint: Use  $[\text{H}^+] = 1 \times 10^{-\text{pH}} \text{ M}$ .**

If a base/acid is diluted with water, the pH will fall/rise to approach 7 (Neutrality) but will not fall below/above 7.

The pH change due to dilution can be calculated using the  $1 \times 10^{-\text{pH}}$  formula to find the change to the  $[\text{H}^+]$ . For a  $10^1$  X dilution, the pH will change by 1 (closer to pH 7) while a 100 X ( $10^2$ ) dilution, the pH will change by 2.

eg. A solution of pH 2.0 is diluted 100 X (from 10 mL to 1000 mL by adding 990 mL of  $\text{H}_2\text{O}$ ). Find its pH.

pH = 2 means the  $[\text{H}^+] = 1.0 \times 10^{-2} \text{ M}$  so for a 100 X dilution, the  $[\text{H}^+] = 1.0 \times 10^{-2} \div 100 = 1.0 \times 10^{-4} \text{ M}$ .

The new pH = -log $_{10}$  $[\text{H}^+] = -\log_{10} 1.0 \times 10^{-4} = 4.0$ . **Hint: change the pH closer to 7 by the  $10^x$  dilution factor.**



The **Mole** (symbol, n and unit, mol) is the **Amount Of Substance** used in calculations: *Hint: amount means mol*

$$\mathbf{1\ CaCl_2(s) + 2\ AgNO_3(aq) \rightarrow 1\ Ca(NO_3)_2(aq) + 2\ AgCl(s).}$$

**Step 2: Find the amount, n(known) using measurements known about a chemical and a mol formula:**

$$n(\text{CaCl}_2) = c \times V = 0.40 \times \frac{250}{1000} = 0.10 \text{ mol}$$

$$n(\text{AgNO}_3) = \frac{m}{M_r} = \frac{17}{170} = 0.10 \text{ mol}.$$

The coefficients in the equation for  $\text{CaCl}_2$  &  $\text{AgNO}_3$  are 1  $\text{CaCl}_2$ : 2  $\text{AgNO}_3$  but the amounts are 0.10: 0.10 mol.

**Step 2a: To identify the chemical in excess, divide each amount (in mol) by its coefficient to view a 1:1 ratio.**

$$\frac{\text{CaCl}_2(\text{s})}{\frac{0.10}{1}} = 0.10 \text{ mol}$$

$$\frac{\text{AgNO}_3(\text{aq})}{\frac{0.10}{2}} = 0.050 \text{ mol}$$

**In Excess (larger value)**

**Limiting Reagent since reaction stops after all is used up**

**Step 2b: If you needed to determine the amount of Excess chemical remaining after the reaction, use:**

**$n(\text{Excess}) = \text{In } 1:1 \text{ ratio } (n(\text{Excess chemical}) - n(\text{Limiting Reagent})) \times \text{Coefficient of Excess chemical}$**

eg.  $n(\text{CaCl}_2)_{\text{XS}} = (0.10 - 0.050) \times 1 = 0.050 \text{ mol}$ . This would leave a concentration of 0.20 M  $\text{CaCl}_2$  ( $\frac{n}{V}$ ).

**Step 3: Use the equation coefficients and n(known chemical NOT in excess) to find the amount, n(Want).**

**$\frac{n(\text{Wanted chem})}{n(\text{Known Chem})} = \frac{\text{Coefficient of Wanted Chem}}{\text{Coefficient of Known Chem}}$  so  $n(\text{Wanted Chem}) = \frac{\text{Coefficient of Wanted Chem}}{\text{Coefficient of Known Chem}} \times$  \_\_\_\_\_**

$$n(\text{AgCl}) = \frac{2}{1} \times n(\text{AgNO}_3) = 1 \times 0.10 = 0.10 \text{ mol}.$$

**Step 4: Convert the amount, n(Want) into the required unit by rearranging a mol formula.**

$$m(\text{AgCl}) = n \times M_r = 0.10 \times (108 + 35.5) = 14.3 \text{ g} = \mathbf{14 \text{ g}}. \quad \text{Use the least number of significant figures.}$$

**Volumetric Analysis involving Stoichiometry (like working out the taste of diluted cordial)**

Stoichiometry is used to determine a sample's concentration in a Volumetric Analysis process called a **Titration**.

**Step 1: A Primary Standard** is accurately weighed & placed it in a 250 mL Volumetric Flask.

Add a suitable solvent ( $\text{H}_2\text{O}$ ) so the bottom of the **meniscus** (U) is on the flask's mark, stopper the flask & shake to dissolve the Primary Standard to form homogenous solution called the **Standard Solution** as its concentration is accurately known. *This is normally done for you.*

Then rinse a **Burette** with the **Standard Solution** and then refill it and record the Initial burette volume (to 2 decimal places). In acid/base titrations the acid is usually in the burette.

**Step 2: Rinse a Pipette** (usually 20 mL) with the **Sample** and refill & transfer this volume called an **Aliquot** to a **Conical Flask**. Add a few drops of a suitable \_\_\_\_\_ if needed.

**Step 3: Begin the Titration** by adding the Standard Solution from the Burette to the sample in the Conical flask and swirl the contents. Stop the Titration when the first hint of a permanent indicator colour change appears called the **End Point** and record the Final burette volume (to 2 decimal places).

This should be very close to the **Equivalence Point** which is when the amounts (n) of the sample and the standard solution are in the same ratio as the coefficients in the balanced equation (so the choice of indicator is important).

eg.  $\boxed{1} \text{ NaOH}(\text{aq}) + \boxed{1} \text{ HCl}(\text{aq}) \rightarrow \text{NaCl}(\text{aq}) + \text{H}_2\text{O}(\text{l})$  Assume  $[\text{HCl}]$  &  $[\text{NaOH}] = 1.0 \text{ M}$  and an acid indicator is used.

Vol1 20 0 pH > 7 due to excess alkaline so the acid indicator (they are weak acids) does not change colour

Vol2 20 20 pH = 7 so amounts (mol) match coefficients = **Equivalence Point** but indicator doesn't change

Vol3 20 20.1 pH < 7 so the acid indicator changes colour = **End Point** (but just past the Equivalence Point).

**Step 4: Use the Final and Initial Burette readings to find the \_\_\_\_\_.**

**Step 5: Repeat the titrations until you achieve 3 concordant titre volumes within 0.1 mL (the difference between the highest and lowest titre volume), then ignore the outliers and find the average titre volume.**

**Step 6: Use Stoichiometry to calculate the [sample].**

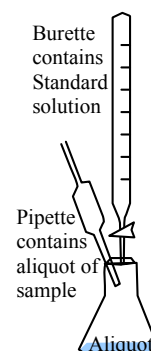
eg. A 20.00 mL aliquot of a sample of NaOH is titrated with 0.100 M HCl. The initial burette volume is 0.30 mL and the indicator changes colour (end point) (final burette volume) at 16.30 mL. Determine the  $[\text{NaOH}]$ .

**Step 1: Write a balanced equation:**  $\text{NaOH}(\text{aq}) + \text{HCl}(\text{aq}) \rightarrow \text{NaCl}(\text{aq}) + \text{H}_2\text{O}(\text{l})$ .

**Step 2: Convert the measurements into mol:**  $n(\text{HCl}) = cV = 0.100 \times (0.01630 - 0.00030 \text{ L}) = 0.0016 \text{ mol}$

**Step 3: Use ratios & coefficients to find n(want):**  $n(\text{NaOH}) = \frac{1}{1} \times n(\text{HCl}) = 0.0016 \text{ mol}$ .

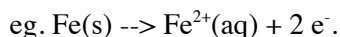
**Step 4: Convert the n(want) into the required unit:**  $[\text{NaOH}] = \frac{n}{V} = \frac{0.0016}{0.020} = 0.080 \text{ M}$ .



**Chapter 16: REDOX CHEMISTRY AND CORROSION***(like metals rusting or silver tarnishing)*

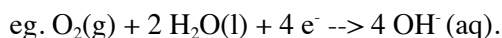
A redox reaction involves an **Oxidation** reaction and a **Reduction** reaction occurring simultaneously. eg. Fe rusting

**Oxidation** involves an \_\_\_\_\_ in positive charge and  $e^-$  are produced on the RHS of the  $\rightarrow$ .



*Hint: Most reactive metals are easily oxidised.*

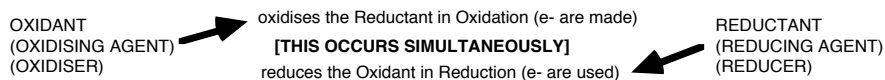
**Reduction** involves a \_\_\_\_\_ in positive charge and  $e^-$  are reacted on the LHS of the  $\rightarrow$ .



*Hint: Most non metals are easily reduced.*

Since the reactions occur simultaneously, the chemical that causes the oxidation of the other chemical is called the

**Oxidant** (eg.  $\text{O}_2$ ) and the chemical that causes the reduction of the other chemical is called the **Reductant** (eg. Fe).



To determine the overall **Redox** equation, the half cell equations are added together so the \_\_\_\_\_ cancel.

If the  $e^-$  do not cancel, each equation is multiplied by a factor so the number of electrons in both half equations are the same and cancel and the total number of atoms on both sides of the reaction arrow must also be equal.

eg. Since the Reduction equation of Oxygen uses  $4e^-$  but the Oxidation reaction of Fe produces only  $2e^-$ , the Fe Oxidation reaction must be  $\times 2$ :  $2\text{Fe(s)} + \text{O}_2(\text{g}) + 2\text{H}_2\text{O(l)} \rightarrow 2\text{Fe}^{2+}(\text{aq}) + 4\text{OH}^-(\text{aq})$  = These produce  $2\text{Fe(OH)}_2(\text{s})$ .

Rusting (Corrosion of metals) is a multi stage process and more Oxygen and Water will react with the Iron

Hydroxide to form Iron Oxide:  $4\text{Fe(OH)}_2(\text{s}) + \text{O}_2(\text{g}) + 2\text{H}_2\text{O(l)} \rightarrow 4\text{Fe(OH)}_3(\text{s}) = 2\text{Fe}_2\text{O}_3 \cdot x\text{H}_2\text{O(l)}$  where  $x = 1 \rightarrow 4$

To prevent Rusting connect a more reactive metal (sacrificial anode) or negative electrical terminal or seal/paint it.

If water is not present, Iron can still undergo Oxidation by Dry Corrosion:  $4\text{Fe(s)} + 3\text{O}_2(\text{g}) \rightarrow 2\text{Fe}_2\text{O}_3(\text{s})$ .

Another method of identifying Oxidation and Reduction in a Redox reaction is to look at the **Oxidation Number**.

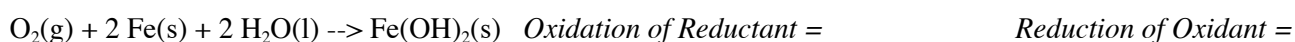
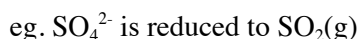
Rule 1: The Oxidation Number (ON) of elements is 0. eg.  $\text{ON}(\text{O}_2) = 0$  and the  $\text{ON}(\text{Fe}) = 0$ .

Rule 2: The Oxidation Number of an ion is equal to its charge but the sign is written first. eg.  $\text{ON}(\text{Fe}^{2+}) = +2$ .

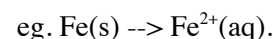
eg. Determine the oxidation of K, Cr and O in  $\text{KMnO}_4$ :  $\text{ON}(\text{K}) =$   $\text{ON}(\text{Mn}) =$   $\text{ON}(\text{O}) =$

In **Oxidation**, the Oxidation Number **increases** while in **Reduction** the Oxidation Number **decreases**.

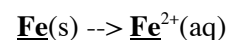
eg. Identify which chemical is undergoing Oxidation and which is undergoing Reduction in the following equation:

**Building Reduction or Oxidation reaction (Half Cell) equations:**

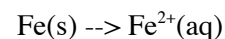
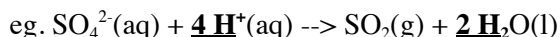
or

**Step 1: Balance the non Oxygen and Hydrogen atoms:**

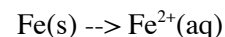
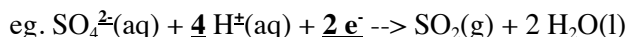
or

**Step 2: Balance the Oxygen atoms using  $\text{H}_2\text{O(l)}$ :**

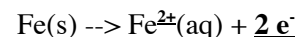
or

**Step 3: Balance the Hydrogen atoms using  $\text{H}^+(\text{aq})$ :**

or

**Step 4: Balance the charges using  $e^-$  by adding the electrons to the more positive side:**

or

**Electrochemical Cells** *(like batteries)*

An **Electrochemical** or **Galvanic Cell** produces electrical energy by a chemical reaction occurring in two half cells.

A half cell contains a conducting metal or C **Electrode** and a conducting solution called an **Electrolyte**.

A **Salt Bridge** (containing  $\text{KNO}_3$  that does not react) connects the two half cells together to complete the circuit and allow the movement of \_\_\_\_\_ between the half cells to neutralize charges.

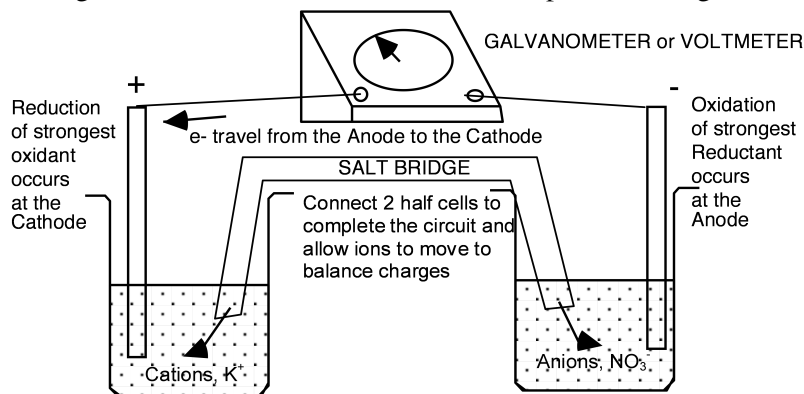
The wires connecting the electrodes to a voltmeter or light make up the **External Circuit** where  $e^-$  move while the Electrolyte where ions move and the Electrodes make up the **Internal Circuit**.

At the **Anode** (-), Oxidation of the strongest reductant occurs to produce  $e^-$  and an \_\_\_\_\_ in positive charge so spectator anions (eg.  $\text{NO}_3^-(\text{aq})$ ) from the salt bridge travel to the anode to balance excess positive charges.

At the **Cathode** (+), \_\_\_\_\_

of the strongest oxidant occurs so  $e^-$  react & there is a decrease in positive charge so spectator cations ( $\text{K}^+(\text{aq})$ ) from the salt bridge travel to the cathode to balance the excess negative charge.

eg. A half cell containing a Zinc anode and  $\text{ZnCl}_2(\text{aq})$  electrolyte is connected to a half cell containing a Copper cathode and  $\text{CuSO}_4(\text{aq})$  electrolyte.



At the Anode: Zinc is oxidized:  $\text{Zn}(\text{s}) \rightarrow \text{Zn}^{2+}(\text{aq}) + 2e^-$ .

The  $\text{Cl}^-$  ions are Spectator ions & ignored.

At the Cathode: Copper ions are reduced:  $\text{Cu}^{2+}(\text{aq}) + 2e^- \rightarrow \text{Cu}(\text{s})$ .

The  $\text{SO}_4^{2-}$  ions are also Spectator ions.

In the overall equation, the electrons in one half cell equation must balance with the electrons in the other half cell equation so they may have to be multiplied by certain factors. eg.  $\text{Zn}(\text{s}) + \text{Cu}^{2+}(\text{aq}) \rightarrow \text{Zn}^{2+}(\text{aq}) + \text{Cu}(\text{s})$ .

The half cells are kept separate to prolong electrical energy production and the cell will continue to produce electricity until there are no more electrolyte ions to react or the electrodes corrode (oxidise).

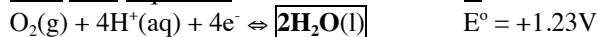
Since  $e^-$  repel from the anode, it is negatively charged while  $e^-$  are attracted to the cathode so it is positively charged.

The **Table of Electrode Potentials** or  **$E^\circ$  Table** shows all chemical types and is used to predict redox reactions.

eg. To predict reactions: Zinc ( $\text{Zn}(\text{s})$ ) in  $\text{ZnCl}_2(\text{aq})$  connected to Iron II Sulfate ( $\text{Fe}^{2+}(\text{aq})$   $\text{SO}_4^{2-}(\text{aq})$ ) with Fe electrode.

**Step I: Use the  $E^\circ$  table to identify ALL species & their reactions. For solutions (aq), include  $\text{H}_2\text{O}$  (1.23, -0.83V).**

#### Half Cell Equation



$\text{SO}_4^{2-} + 4\text{H}^+ + 2e^- \rightleftharpoons \text{H}_2\text{SO}_3 + \text{H}_2\text{O} \quad E^\circ = +0.20\text{V}$  but since the solution is not acidified ( $\text{H}^+$ ) then **not all the reactants are present** so the reaction **cannot** occur.

Be careful identifying the species present & their states. eg. In  $\text{FeCl}_2$  there is  $\text{Fe}^{2+}$ , 2  $\text{Cl}^-$  not  $\text{Cl}_2(\text{g})$ .

**Step II: Select the strongest Oxidant on the LHS with the highest  $E^\circ$  that undergoes Reduction at the cathode. The cathode itself cannot react.**

\_\_\_\_\_ on the left side of the above reduction reactions **high** on the  $E^\circ$  table are stronger and prefer to be reduced at the Cathode than oxidants lower on the table.

#### REDUCTION

( $\uparrow E^\circ$  on LHS):

**Step III: Select the strongest Reductant on the**

**RHS with the lowest  $E^\circ$  that undergoes**

**Oxidation at the anode (that is written as a**

**Reduction reaction in the table so it needs to be reversed)**

| Electrode Potential Table   |  | $E^\circ$<br>(V) |
|---|--|------------------|
| Oxidant (Reduction) + $e^- \rightleftharpoons$ Reductant (Oxidation)  |  |                  |
| $\text{F}_2(\text{g}) + 2e^- \rightleftharpoons 2\text{F}^-(\text{aq})$   |  | +2.87            |
| $\text{PbO}_2(\text{s}) + \text{HSO}_4^-(\text{aq}) + 3\text{H}^+(\text{aq}) + 2e^- \rightleftharpoons \text{PbSO}_4(\text{s}) + 2\text{H}_2\text{O}(\text{l})$ |  | +1.69            |
| $\text{MnO}_4^-(\text{aq}) + 8\text{H}^+(\text{aq}) + 5e^- \rightleftharpoons \text{Mn}^{2+}(\text{aq}) + 4\text{H}_2\text{O}(\text{l})$                        |  | +1.52            |
| $\text{Au}^+(\text{aq}) + e^- \rightleftharpoons \text{Au}(\text{s})$   |  | +1.50            |
| $\text{Cl}_2(\text{g}) + 2e^- \rightleftharpoons 2\text{Cl}^-(\text{aq})$   |  | +1.36            |
| $\text{O}_2(\text{g}) + 4\text{H}^+(\text{aq}) + 4e^- \rightleftharpoons 2\text{H}_2\text{O}(\text{l})$   |  | +1.23            |
| $\text{Br}_2(\text{g}) + 2e^- \rightleftharpoons 2\text{Br}^-(\text{aq})$   |  | +1.08            |
| $\text{Ag}^+(\text{aq}) + e^- \rightleftharpoons \text{Ag}(\text{s})$   |  | +0.80            |
| $\text{Fe}^{3+}(\text{aq}) + e^- \rightleftharpoons \text{Fe}^{2+}(\text{aq})$  |  | +0.77            |
| $\text{I}_2(\text{aq}) + 2e^- \rightleftharpoons 2\text{I}^-(\text{aq})$  |  | +0.54            |
| $\text{O}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l}) + 4e^- \rightleftharpoons 4\text{OH}^-(\text{aq})$  |  | +0.40            |
| $\text{Cu}^{2+}(\text{aq}) + 2e^- \rightleftharpoons \text{Cu}(\text{s})$   |  | +0.34            |
| $\text{SO}_4^{2-}(\text{aq}) + 4\text{H}^+(\text{aq}) + 2e^- \rightleftharpoons \text{H}_2\text{SO}_3(\text{aq}) + \text{H}_2\text{O}(\text{l})$                |  | +0.20            |
| $\text{Sn}^{4+}(\text{aq}) + 2e^- \rightleftharpoons \text{Sn}^{2+}(\text{aq})$   |  | +0.15            |
| $\text{S}(\text{s}) + 2\text{H}^+(\text{aq}) + 2e^- \rightleftharpoons \text{H}_2\text{S}(\text{g})$  |  | +0.14            |
| $2\text{H}^+(\text{aq}) + 2e^- \rightleftharpoons \text{H}_2(\text{g})$   |  | 0.00             |
| $\text{Pb}^{2+}(\text{aq}) + 2e^- \rightleftharpoons \text{Pb}(\text{s})$   |  | -0.13            |
| $\text{Sn}^{2+}(\text{aq}) + 2e^- \rightleftharpoons \text{Sn}(\text{s})$   |  | -0.14            |
| $\text{Ni}^{2+}(\text{aq}) + 2e^- \rightleftharpoons \text{Ni}(\text{s})$   |  | -0.25            |
| $\text{PbSO}_4(\text{s}) + \text{H}^+(\text{aq}) + 2e^- \rightleftharpoons \text{Pb}(\text{s}) + \text{HSO}_4^-(\text{aq})$                                     |  | -0.36            |
| $\text{Fe}^{2+}(\text{aq}) + 2e^- \rightleftharpoons \text{Fe}(\text{s})$   |  | -0.44            |
| $\text{Zn}^{2+}(\text{aq}) + 2e^- \rightleftharpoons \text{Zn}(\text{s})$   |  | -0.76            |
| $2\text{H}_2\text{O}(\text{l}) + 2e^- \rightleftharpoons \text{H}_2(\text{g}) + 2\text{OH}^-(\text{aq})$  |  | -0.83            |
| $\text{Mn}^{2+}(\text{aq}) + 2e^- \rightleftharpoons \text{Mn}(\text{s})$   |  | -1.03            |
| $\text{Al}^{3+}(\text{aq}) + 3e^- \rightleftharpoons \text{Al}(\text{s})$   |  | -1.67            |
| $\text{Mg}^{2+}(\text{aq}) + 2e^- \rightleftharpoons \text{Mg}(\text{s})$   |  | -2.34            |
| $\text{Na}^+(\text{aq}) + e^- \rightleftharpoons \text{Na}(\text{s})$   |  | -2.71            |
| $\text{Ca}^{2+}(\text{aq}) + 2e^- \rightleftharpoons \text{Ca}(\text{s})$   |  | -2.87            |
| $\text{K}^+(\text{aq}) + e^- \rightleftharpoons \text{K}(\text{s})$   |  | -2.93            |
| $\text{Li}^+(\text{aq}) + e^- \rightleftharpoons \text{Li}(\text{s})$   |  | -3.02            |

Be careful  
as there are  
2  $\text{H}_2\text{O}$ ,  $\text{Fe}^{2+}$   
&  $\text{Sn}^{2+}$   
equations

Increasing Oxidant (Reduction) Strength

Increasing Reductant (Oxidation) Strength

Reductants on the right side of the above reduction reactions **low** on the  $E^\circ$  table are stronger & prefer to be oxidised at the Anode more than reductants higher on the table. For Oxidation, **reverse** the Reduction equations in the table.

## OXIDATION

( $\downarrow E^\circ$  on RHS):

**Step IV: For a Galvanic Cell reaction to occur it MUST obey the following rules:**

- the Oxidant's Reduction must have a **HIGHER  $E^\circ$**  than the Reductant's Oxidation (Exothermic profile).
  - the Reduction and Oxidation reactions must occur in separate half cells and not in the same half cell.
  - the  $\frac{1}{2}$  cell reactions with the greatest positive difference in  $E^\circ$  are favoured as they release the most energy.
- The overall equation must include states but no  $e^-$  so multiply each  $\frac{1}{2}$  reaction by a factor and add them:

## OVERALL

(with no  $e^-$ ):

**Step V: Calculate the Electromotive Force (EMF) or Potential Difference (PD) produced by the redox reaction.**

**$= E^\circ(\text{Reduction at cathode}) - E^\circ(\text{Oxidation at anode})$ . The  $-$  is due to the Oxidation reaction being reversed.**

**It must be GREATER than 0 for the cell's redox reaction to produce electrical energy (exothermic).**

Cell EMF:

**Summary comparing Galvanic Cells to Electrolytic cells (Electrolytic cells are introduced in Unit 3/4):**

| Characteristics                      | Electrochemical Cell  | Electrolytic Cell                        |
|--------------------------------------|---|--|
| Cell setup                           |   |  |
| Anode charge                         | - since electrons naturally repelled/leave here   | + since power forces electrons to leave  |
| Choice of Oxidation                  | Lowest $E^\circ$ oxidant on right side of $E^\circ$ table                                   |  |
| Cathode charge                       | + since electrons naturally attracted here  | - since power forces electrons here      |
| Choice of Reduction                  | Highest $E^\circ$ oxidant on left side of $E^\circ$ table                                   |  |
| Electron flow                        | From anode to cathode   |  |
| Flow of ions to prevent polarisation | Spectator anions from salt bridge travel to anode & spectator cations travel to the cathode |  |
| Cell EMF (voltage)                   | Greater than 0 so reaction is spontaneous   | Less than 0 so apply power for reaction  |
| Energy conversions                   | Chemical to Electrical = Exothermic ( $\neg \neg$ )   | Electrical to Chemical = Endothermic     |
| Uses                                 | Batteries   | Plating metal jewellery, refining metals |

An **Activity Series** shows the reactivity of Metals undergoing Oxidation (corrosion or rusting) in order of least reactive to most reactive: Au, Pt, (C), Ag, ( $O_2$ ,  $H_2O$ ), Cu, Pb, Sn, Ni, Fe, Mn, Cr, Zn, ( $H_2O$ ), Al, Mg, Na, Ca, K, Li.

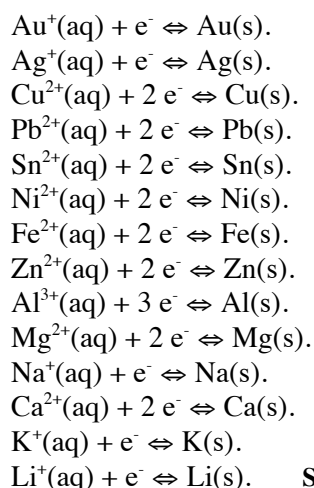
**Strongest Oxidant prefers Reduction -->**

**Weakest Reductant**

Reduction  
of Cation  
to form  
Metal  
element  
at the  
Cathode

Oxidation  
of Metal  
element to  
form  
Cation at  
the Anode

**Weakest Oxidant**



**Strongest Reductant prefers Oxidation <--**

Metals before  $O_2$ ,  $H_2O$  will not undergo \_\_\_\_\_ readily in Oxygen & Water (but there are some exceptions like Silver, Ag that tarnishes in air,  $O_2$ ). However Gold, Au and Platinum, Pt never oxidise in air.

Metals between  $O_2$ ,  $H_2O$  and  $H_2O$  will undergo Oxidation (corrode) in the presence of both  $O_2$  and  $H_2O$ . eg. Pb, Fe. Any Metal by itself after  $H_2O$  is oxidised in water since they are weaker reductants than  $H_2O$ . eg. Na, K.

A reaction only occurs if the Oxidant (LHS) is \_\_\_\_\_ than the Reductant (RHS) on the **Activity Series**.

A related Reductant and Oxidant which differ by the  $e^-$  form a **Conjugate Redox Pair**. eg.  $Li^+/Li$  or  $Cl_2/Cl^-$ .

Lithium, Li is the strongest reductant and more likely to undergo Oxidation so  $Li^+$  is the weakest oxidant.

The cation of Gold,  $Au^+$  is a strong oxidant and will undergo Reduction so Au is a \_\_\_\_\_ reductant.

Metals in the middle of the series can behave either as an oxidant or reductant depending upon the position on the series of the other chemical reacting. eg. Cu reacts with  $Ag^+$  or  $Cu^{2+}$  reacts with Zn but Cu can't react with  $Zn^{2+}$ .

Compare with non metals that are strong oxidants that prefer reduction to form anions & seldom undergo oxidation.

### Chapter 17: GREEN CHEMISTRY (like trying to rotect us from pollutants & poisons)

While Environmental Chemistry tries to reduce waste/danger/pollutants AFTER production (Secondary prevention),

**Green Chemistry** uses 12 Principles to reduce waste/danger/pollutants DURING production (Primary prevention):

- |   |   |
|---|---|
| a). Prevent waste instead of having to clean up.          | b). Design safer (non-toxic) products.                    |
| c). Design _____ hazardous chemical processes.            | d). Design for degradation (breakdown).                   |
| e). Use renewable (plant not fossil fuel) raw materials.  | f). Avoid chemical derivatives (temporary chemicals).     |
| g). Maximise atom economy (product has most atoms).       | h). Use safer solvents & reaction conditions.             |
| i). Minimise energy use for reactions (use at room temp). | j). Use catalysts to avoid using large reactants amounts. |
| k). Analyse pollutants in real time (during reactions).   | l). Minimise accidents & pollution risks.                 |

Atom economy (fewest byproducts & pathway stages) =  $\frac{Mr(\text{Product}) \times \text{coeff} \times 100}{\text{sum}(Mr \text{ of Reactants} \times \text{coeff})}$ . Product yield =  $\frac{\text{Actual product mass} \times 100}{\text{Calculated product mass}}$ .

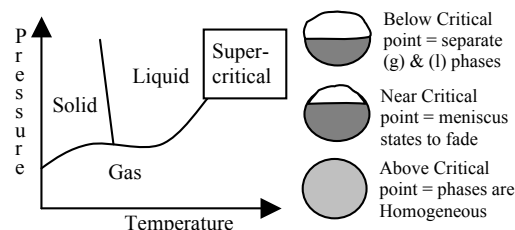
eg. Using supercritical Carbon Dioxide produced by condensing  $CO_2$  to 73 atm & heating it above  $31.1^\circ C$  to replace flammable Organic (Carbon based) solvents to decaffeinate coffee, extract fragrances & in dry cleaning clothing.

It is a **Supercritical Fluid** since it exists above its **Critical Point** (the

\_\_\_\_\_ temperature & pressure at which a substance exists as a vapour and liquid (where it boils & condenses)) so its properties are between a gas and liquid. Non polar organic solvents are soluble in supercritical  $CO_2$  but polar/ionic substances are insoluble.

eg. Supercritical Water is less polar due to the movement of H atoms at high pressure and temperature (221 atm &  $374^\circ C$ ) so it can be used to separate the moisture content from the Biomass vapour fuel in biowaste/compost etc.

eg. Ionic liquids are liquids at room temperature and normally have a large metal cation bonded to a small non metal anion. They can dissolve ionic chemicals and some organic chemicals and polymers.

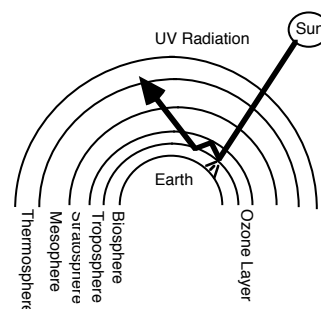


### Chapter 18 & 19: THE ATMOSPHERE & ENVIRONMENTAL ISSUES

The small layer of the earth's crust that supports life is called the **Biosphere**.

In the Stratosphere, ozone ( $O_3$ ) in concentrations of up to 10 ppm (parts per million) absorbs most of the UV radiation produced by the sun. The formation of ozone requires UV radiation to split a molecule of oxygen to produce **Monoatomic** oxygen:  $O_2 \rightarrow O + O$  followed by  $O_2 + O \rightarrow O_3$  in the presence of \_\_\_\_\_ radiation. These atoms combine with other oxygen molecules to form ozone if a third molecule such as  $N_2$  is available to absorb the heat energy released:  $O + N_2 + O_2 \rightarrow O_3 + N_2$ .

Gases such as ChloroFluorocarbons (CFC's. eg  $CCl_3F$ ), Nitrous Oxide ( $NO$ ) & Halons contribute to the **Greenhouse Effect** (atmosphere layers trap warm air that lead to rising sea levels), ecological changes (changes to the environment) and climatic changes) & also destroy the ozone layer (which prevents UV radiation reaching the earth's surface):  $NO(g) + O_3(g) \rightarrow NO_2(g) + O_2(g)$  and then  $NO_2(g) + O(g) \rightarrow NO(g) + O_2(g)$  as well as  $CCl_3F(g) \rightarrow CCl_2F(g) + Cl(g)$ , then



$\text{Cl(g)} + \text{O}_3\text{(g)} \rightarrow \text{ClO(g)} + \text{O}_2\text{(g)}$  and then  $\text{ClO(g)} + \text{O(g)} \rightarrow \text{O}_2\text{(g)} + \text{Cl(g)}$ .

The overall reaction for the ozone breakdown is  $\text{O}_3\text{(g)} + \text{O(g)} \rightarrow 2 \text{O}_2\text{(g)}$

UV radiation catalyses ozone production and breakdown in a never ending cycle.

In 2006, Australia attended the Asia-Pacific Partnership on Clean Development & Climate (AP6) non treaty meetings to reduce Greenhouse emissions (by an amount they state), back to 1990 levels.

The **Kyoto Protocol** states it is mandatory to reduce Greenhouse gas (Carbon Dioxide, Nitrous Oxide, Methane, Sulfur Hexafluoride, Hydrofluorocarbons (HFCs) & Perfluorocarbons (PFCs)) emissions by 29% (est. 2010 level). Australia can increase their levels by 8% while other countries must reduce their levels but Australia has not signed.

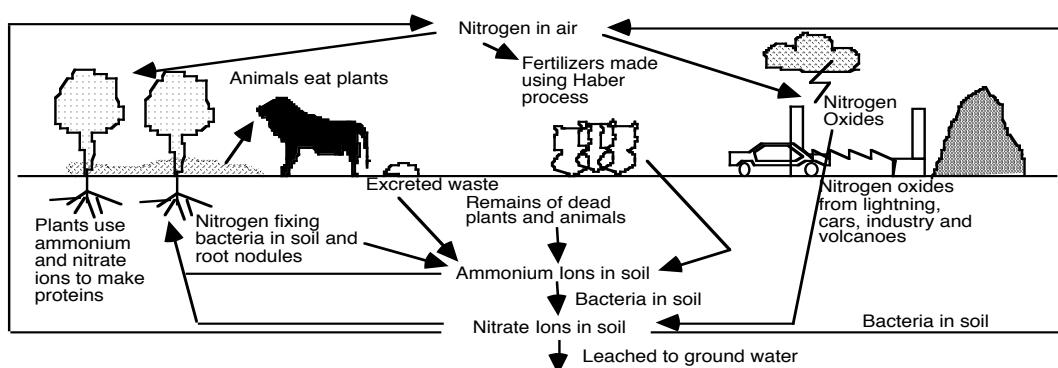
Oxygen is produced by **Photosynthesis** in plants but requires light to strike Chlorophyll in leaves to catalyse the reaction. Oxygen is used up in the reverse reaction called \_\_\_\_\_ that occurs in plant and animal cells

during the day and night:  $\text{C}_6\text{H}_{12}\text{O}_6\text{(aq)} + 6 \text{O}_2\text{(g)} \xrightleftharpoons[\text{Photosynthesis}]{\text{Respiration}} 6 \text{CO}_2\text{(g)} + 6 \text{H}_2\text{O(l)} + \text{Energy}$

Glucose ( $\text{C}_6\text{H}_{12}\text{O}_6$ ) polymerises to form the complex sugars starch (animal/plant) & cellulose (plant structure).

**Matter Cycles:** The cycle of an element can be shown using a **Matter Cycle**. (like recycling)

**Nitrogen Cycle:** The amount of nitrogen in the cycle is fairly constant.



**Nitrogen Fixation** converts unreactive  $\text{N}_2$  gas (due to the triple covalent bond) into soluble (usable) nitrogen

nutrients: eg. Plants use ammonium ions,  $\text{NH}_4^+$  to form amino acids & simple proteins which are eaten by animals.

a). Bacteria in the soil & legume plant root nodules have enzymes to catalyse the reduction of  $\text{N}_2$  (oxidation no. = 0) into  $\text{NH}_3$  (oxidation no. = -3) or  $\text{NH}_4^+$  (oxidation no. = \_\_\_\_\_) so they can be absorbed into plants.

b). In industry, the Haber process reduces  $\text{N}_2$  to ammonia with a Fe catalyst at  $500^\circ\text{C}$ :  $\text{N}_2\text{(g)} + 3 \text{H}_2\text{(g)} \rightleftharpoons 2 \text{NH}_3\text{(g)}$ .

c). In **Nitrification**, soil bacteria oxidise Nitrogen in ammonia or ammonium ions,  $\text{NH}_4^+$  into nitrate ions,  $\text{NO}_3^-$  (oxidation no. = -3  $\rightarrow$  \_\_\_\_\_) which are soluble in water and leach from the soil into the ground water.

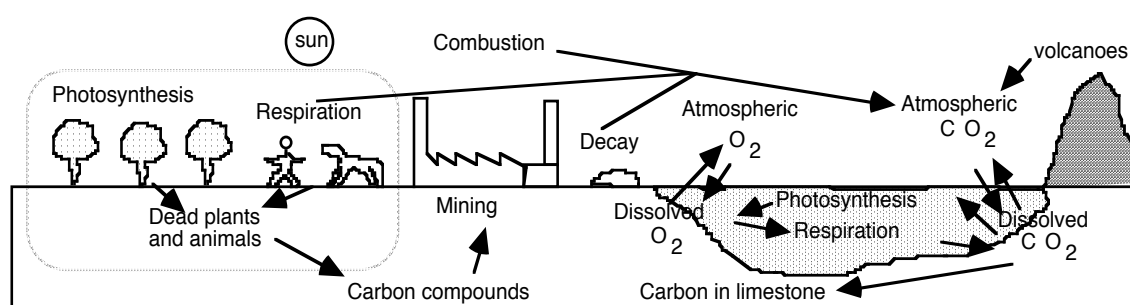
d). Engines, lightning, volcanoes & industry burn nitrogen forming clear, water insoluble Nitrogen monoxide (NO) that combines with oxygen to form brown, toxic Nitrogen Dioxide ( $\text{NO}_2$ ) that is soluble & forms dilute Nitric acid ( $\text{HNO}_3$ ) & Nitrous acid ( $\text{HNO}_2$ ) that can be used to make fertilizers.  $\text{N}_2\text{(g)} + \text{O}_2\text{(g)} \rightarrow 2 \text{NO(g)}$ , then  $2 \text{NO(g)} + \text{O}_2\text{(g)} \rightarrow 2 \text{NO}_2\text{(g)}$  which dissolves in rain:  $2 \text{NO}_2\text{(g)} + \text{H}_2\text{O(l)} \rightarrow \text{HNO}_3\text{(aq)} + \text{HNO}_2\text{(aq)}$ .

In **Denitrification**, certain bacteria reduce Nitrogen in Nitrates back to Nitrogen (oxidation no. = +5  $\rightarrow$  0).

Animals excrete waste Nitrogen as urea,  $\text{CO}(\text{NH}_2)_2$ . When plants and animals decay, they release  $\text{NH}_4^+$  and  $\text{N}_2$ .

Farmers use ammonium fertilizers and manure & plant legumes (organic) to supply fixed nitrogen to plants. These fertilizers normally contain compounds of Nitrogen:  $\text{NH}_4\text{NO}_3$ ,  $(\text{NH}_4)_2\text{SO}_4$  (from ammonia &  $\text{H}_2\text{SO}_4$ ).

**Carbon-Oxygen Cycle:** The amount of Carbon & Oxygen in the environment remains fairly constant.



Some carbon is kept in the ground to form fossil fuels such as oil or coal which can be used in **Combustion** reactions to produce energy ( $\text{C(s)} + \text{O}_2\text{(g)} \rightarrow \text{CO}_2\text{(g)}$ ) or forms oxides ( $\text{CO}_3^{2-}\text{(aq)}$ ) which combine with metal ions to form compounds ( $\text{Ca}^{2+}\text{(aq)} + \text{CO}_3^{2-}\text{(aq)} \rightarrow \text{CaCO}_3\text{(s)}$ ), Calcium Carbonate (limestone or marble).

Carbon dioxide is slightly soluble in water:  $\text{H}_2\text{O(l)} + \text{CO}_2\text{(g)} \rightleftharpoons \text{H}_2\text{CO}_3\text{(aq)}$  unstable  $\rightleftharpoons \text{HCO}_3^-\text{(aq)} + \text{H}^+\text{(aq)}$  and is released when dead animals and plants decay.  $\text{O}_2$  is consumed &  $\text{CO}_2$  is produced by \_\_\_\_\_ in animals & plants during the day & night while  $\text{O}_2$  is produced &  $\text{CO}_2$  is removed from the atmosphere by

**Photosynthesis** in plant leaves during the day:  $\text{C}_6\text{H}_{12}\text{O}_6\text{(aq)} + 6 \text{O}_2\text{(g)} \xrightarrow{\text{Photosynthesis}} 6 \text{CO}_2\text{(g)} + 6 \text{H}_2\text{O(l)}$ .

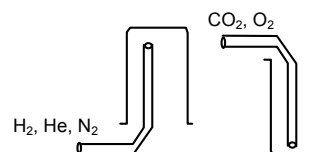
The amount of Oxygen formed by Photosynthesis is \_\_\_\_\_ than that used in plant Respiration.

The rising levels of  $\text{CO}_2$  &  $\text{CH}_4$  (methane) contribute towards the atmospheric warming (Greenhouse Effect).

### Chapter 20: GASES OF THE ATMOSPHERE.

The atmosphere contains 20%  $\text{O}_2$ , 0.035 %  $\text{CO}_2$ , 79 %  $\text{N}_2$  and small amounts of Noble Gases (He and Ne).

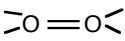
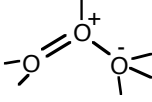
The Molecular Mass,  $M_r$  of Air is approx. 31, so lighter gases can be collected by the downward displacement of air in inverted beakers/tubes include CO ( $M_r = 28$ ),  $\text{N}_2$  ( $M_r = 28$ ),  $\text{CH}_4$  ( $M_r = 16$ ), He ( $A_r = 4$ ) and  $\text{H}_2$  ( $M_r = 2$ ), while heavier gases can be collected by the upward displacement of air in beakers/tubes include  $\text{CO}_2$  ( $M_r = 44$ ) and  $\text{O}_2$  ( $M_r = 32$ ).



Gases such as  $\text{O}_2$  and  $\text{CO}_2$  are less soluble in water than most ionic solids. Gas solubility \_\_\_\_\_ as the temperature increases but is increased by forcing the gas into solution with pressure. eg.  $\text{CO}_2$  in Soft drinks.

| Gas   | Characteristics   | Sources  |
|---|---|--|
| Carbon Dioxide<br>$\text{>O=C=O<}$<br><br>The amount of $\text{CO}_2$ released by engines can be controlled tuning the fuel/air mixture and recycling incomplete combusted gases which also lowers engine temp. | It is a colorless, odourless gas. When cooled & condensed it will not liquefy but solidifies to <b>Dry Ice</b> at $-78^\circ\text{C}$ .<br><br>It will <b>sublime</b> when heated, turning from solid directly to gas.<br><br>It is slightly soluble in water but the product Carbonic acid, $\text{CO}_2\text{(g)} + \text{H}_2\text{O(l)} \rightarrow \text{H}_2\text{CO}_3\text{(aq)}$ is a very unstable weak acid and ionises:<br>$\text{H}_2\text{CO}_3\text{(aq)} + \text{H}_2\text{O(l)} \rightleftharpoons \text{H}_3\text{O}^+\text{(aq)} + \text{HCO}_3^-\text{(aq)}$ .<br>Soft Drinks are carbonated by blowing $\text{CO}_2$ into the soft drink – making it slightly acidic. The gas is more soluble when chilled.<br><br>It is detected by extinguishing flames and by blowing the gas through Limewater ( $\text{Ca(OH)}_2$ ), causing it to turn milky due to the formation of a Calcium Carbonate precipitate:<br>$\text{Ca(OH)}_2\text{(aq)} + \text{CO}_2\text{(g)} \rightarrow \text{CaCO}_3\text{(s)} + \text{H}_2\text{O(l)}$<br><br>If more is blown through, the precipitate of Calcium Carbonate will _____:<br>$\text{CaCO}_3\text{(s)} + \text{H}_2\text{O(l)} + \text{CO}_2\text{(g)} \rightarrow \text{Ca(HCO}_3)_2\text{(aq)}$ | Produced in decomposition reactions of Metal Carbonates with heat:<br>$\text{CaCO}_3\text{(s)} \rightarrow \text{CaO(s)} + \text{CO}_2\text{(g)}$ .<br><br>Produced by reacting Metal Carbonates with acids:<br>$\text{CaCO}_3\text{(s)} + 2 \text{HCl(aq)} \rightarrow \text{CaCl}_2\text{(aq)} + \text{H}_2\text{O(l)} + \text{CO}_2\text{(g)}$<br><br>Produced in Aerobic Respiration of Glucose in plants and animals:<br>$\text{C}_6\text{H}_{12}\text{O}_6\text{(aq)} + 6 \text{O}_2\text{(g)} \rightarrow 6 \text{CO}_2\text{(g)} + 6 \text{H}_2\text{O(l)} + \text{Energy}$<br><br>It's also produced in Anaerobic Respiration (Fermentation) using a yeast catalyst:<br>$\text{C}_6\text{H}_{12}\text{O}_6\text{(aq)} \rightarrow 2 \text{C}_2\text{H}_5\text{OH(aq)} + 2 \text{CO}_2\text{(g)}$<br><br>Produced in _____ reactions in an excess of air:<br>$\text{C}_3\text{H}_8\text{(g)} + 5 \text{O}_2\text{(g)} \rightarrow 3 \text{CO}_2\text{(g)} + 4 \text{H}_2\text{O(l)}$ |
| Carbon Monoxide, CO<br>$-\text{O}^+ \equiv \text{C}^-$  | It is a colourless, odourless gas that is highly poisonous as it competes against oxygen to bind to haemoglobin in blood to produce carboxyhaemoglobin, not oxyhaemoglobin.   | Carbon Monoxide is produced by the combustion of fuels in limited oxygen:<br>$\text{C}_3\text{H}_8\text{(g)} + \frac{7}{2} \text{O}_2\text{(g)} \rightarrow 3 \text{CO(g)} + 4 \text{H}_2\text{O(l)}$<br><br>Engines use catalytic converters to convert CO to $\text{CO}_2$ : $2 \text{CO(g)} + \text{O}_2\text{(g)} \rightarrow 2 \text{CO}_2\text{(g)}$ .   |



| Gas   | Characteristics   | Sources  |
|---|---|--|
| Oxygen, O <sub>2</sub><br><br>(Diatomic molecule - contains 2 atoms) | It is a colourless, odourless gas which is slightly soluble in water and reacts with most metals causing them to Corrode (rust)<br>$(4 \text{ Fe(s)} + 3 \text{ O}_2\text{(g)} \rightarrow 2 \text{ Fe}_2\text{O}_3\text{(s)})$ and supports combustion $(\text{C(s)} + \text{O}_2\text{(g)} \rightarrow \text{CO}_2\text{(g)})$ or<br>$\text{C}_x\text{H}_y\text{(g)} + (x + \frac{y}{4})\text{O}_2\text{(g)} \rightarrow x\text{CO}_2\text{(g)} + \frac{y}{2}\text{H}_2\text{O(l)}.$<br>Its presence can be detected by observing if a glowing splint ignites into flame.   | Produced by Fractional Distillation of Liquid Air in which the air is chilled and allowed to warm slowly so Oxygen boils at a specific temperature and is collected.<br><br>In nature it is produced in _____ occurring in plants (including algae).<br>Also produced by the decomposition of Hydrogen Peroxide using MnO <sub>2</sub> catalyst:<br>$2 \text{ H}_2\text{O}_2\text{(aq)} \xrightarrow{\text{MnO}_2} 2 \text{ H}_2\text{O(l)} + \text{O}_2\text{(g)}$  |
| Ozone, O <sub>3</sub><br><br>(Triatomic)                             | It is very reactive, has a sweet smell but is an irritant, destroys viruses and bacteria and breaks double Carbon Carbon bonds in rubber.<br>It decomposes readily. eg. $2 \text{ O}_3\text{(g)} \rightarrow 3 \text{ O}_2\text{(g)}.$  | Produced by the action of high energy (electricity or lightning or UV radiation) on O <sub>2</sub> : $3 \text{ O}_2\text{(g)} \rightarrow 2 \text{ O}_3\text{(g)}.$  |
| Nitrogen, N <sub>2</sub><br>$\text{-- N} \equiv \text{N --}$  | It is a colourless, odourless, water insoluble gas that is fairly unreactive due to the triple bond.<br>However, at high temperatures, it will burn:<br>$\text{N}_2\text{(g)} + \text{O}_2\text{(g)} \rightarrow 2 \text{ NO(g)},$ Nitrous Oxide<br>Nitrogen is used in fertilizers, (NH <sub>4</sub> ) <sub>2</sub> SO <sub>4</sub> or NH <sub>4</sub> NO <sub>3</sub> and as ammonia, NH <sub>3</sub> in cleaners.  | Produced by Fractional Distillation of Liquid Air.<br>N <sub>2</sub> is released in Denitrification of Nitrates (NO <sub>3</sub> <sup>-</sup> (aq)) in plants by specific bacteria.<br><br>Other soil bacteria can convert N <sub>2</sub> into NH <sub>4</sub> <sup>+</sup> in a process called _____.   |
| Nitrous Oxide, NO and Nitrogen Dioxide, NO <sub>2</sub>   | Nitrogen Monoxide is a colourless gas that reacts in air & is insoluble in water & is collected by the downward displacement of water in a test tube.<br>Nitrogen Monoxide, NO is reactive in air:<br>$2\text{NO(g)} + \text{O}_2\text{(g)} \rightarrow 2\text{NO}_2\text{(g)}.$<br>Nitrogen Dioxide, NO <sub>2</sub> is a brown gas that is toxic & water soluble and dissolves to form two acids:<br>$2 \text{ NO}_2\text{(g)} + \text{H}_2\text{O(l)} \rightarrow \text{HNO}_3\text{(aq)} + \text{HNO}_2\text{(aq)}.$<br><div style="text-align: center;">Nitric acid    Nitrous acid</div> These acids react with certain hydrocarbons and ozone to produce Peroxyacetyl Nitrates (PAN) which is the major component of <b>Photochemical Smog</b> . | Nitrogen Monoxide (laughing gas) can be produced by the reaction between Copper and 50 % Nitric acid:<br>$3 \text{ Cu(s)} + 8 \text{ HNO}_3\text{(aq)} \rightarrow 3 \text{ Cu(NO}_3)_2\text{(aq)} + 4 \text{ H}_2\text{O(l)} + 2 \text{ NO(g)}$ or by burning Nitrogen in air:<br>$\text{N}_2\text{(g)} + \text{O}_2\text{(g)} \rightarrow 2 \text{ NO(g)}.$<br>Nitrogen Dioxide can be produced by the reaction between Copper & concentrated Nitric acid: $\text{Cu(s)} + 4 \text{ HNO}_3\text{(aq)} \rightarrow \text{Cu(NO}_3)_2\text{(aq)} + 2 \text{ H}_2\text{O(l)} + 2 \text{ NO}_2\text{(g)}.$ |
| Methane, CH <sub>4</sub>  | It is a colourless, odourless, water insoluble gas. Present in Natural gas and used as a fuel since it readily undergoes combustion or can also be used in fuel cells to produce electrical energy:<br>$\text{CH}_4\text{(g)} + 2 \text{ O}_2\text{(g)} \rightarrow \text{CO}_2\text{(g)} + 2 \text{ H}_2\text{O(g)}.$  | Present in Natural Gas, fractions of Crude Oil and produced naturally in decay process of biological matter which can be collected and used as a fuel (Biogas). It is also Greenhouse Gas.   |
| Noble or Inert Gases of Group VIII (Monoatomic molecule)  | They are colourless, odourless and unreactive due to the atom's full _____.<br>Helium is used in balloons since it is lighter than air, Neon and Xenon are used as a gas in light bulbs and Argon is used in welding.   | Produced by Fractional Distillation of _____ where air is chilled and allowed to warm slowly so the Noble Gases each boil at a specific temperature and are collected.   |

| Gas  | Characteristics   | Sources   |
|--|---|---|
| Sulfur Dioxide, SO <sub>2</sub> and Sulfur Trioxide, SO <sub>3</sub> | Both are colourless gases that are soluble in water and produce acidic solutions:<br>$\text{SO}_2(\text{g}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{H}_2\text{SO}_3(\text{aq})$ , Sulfurous acid<br>$\text{SO}_3(\text{g}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{H}_2\text{SO}_4(\text{aq})$ , Sulfuric acid<br>These acids are responsible for _____ | They are produced by burning Sulfur, S (from crude oil impurities) in air (O <sub>2</sub> ).<br>$\text{S}(\text{s}) + \text{O}_2(\text{g}) \rightarrow \text{SO}_2(\text{g})$ , Sulfur Dioxide.<br>$2 \text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2 \text{SO}_3(\text{g})$ ,<br>Sulfur Trioxide |

### Chapter 21: PHYSICAL PROPERTIES OF GASES

The **Kinetic Molecular Theory Of Gases** explains that the volume of the gas molecules is smaller than the volume they occupy, gas particles move in a \_\_\_\_\_ line motion in many directions and the forces between the particles are very weak, so when the particles collide, they collide **Elastically** (without energy loss).

The average kinetic energy (energy due to the motion of particles) increases with temperature but at any time, the kinetic energy of most gas particles can vary. As temperature increases, the solubility of gases decreases.

This Theory explains why a smell travels through a room by the process called **Diffusion** in which low Molecular Mass gas molecules move faster throughout the container than higher Molecular Mass gas molecules and why changing the size of a container is the only way to change the mass of a gas since its volume will be different.

If an "Ideal" gas could exist, there would be no forces between the gas molecules so the gas could never be condensed into a liquid even at absolute zero (\_\_\_\_\_°C).

Due to the spaces between gas particles, gases can be compressed causing a **Pressure** force ( $\text{Pressure} = \frac{\text{Force (F in N)}}{\text{Area (A in m}^2\text{)}}$ )

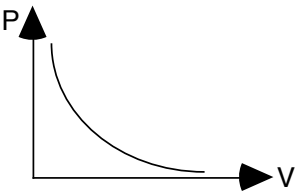
to build up against the walls of the container as the particles still collide with one another & the container. Pressure is measured using a **Barometer**. Each gas particle contributes its own **Partial Pressure** towards the pressure of the container, due to the amount of a particular gas particle present & its \_\_\_\_\_ (since their weight is a force).

e.g. If  $p(\text{O}_2) = 20 \text{ kPa}$ ,  $p(\text{CO}_2) = 30 \text{ kPa}$ ,  $p(\text{N}_2) = 90 \text{ kPa}$  then  $p(\text{air mixture of O}_2, \text{CO}_2 \text{ \& N}_2) = 20+30+90 = 140 \text{ kPa}$ .

The conversion of units is  $760 \text{ mmHg} = 1 \text{ atmosphere} = 101325 \text{ Pa (or Nm}^{-2}\text{)} = 101.325 \text{ kPa} = 1.013 \text{ Bar}$

$$\text{mmHg} \begin{matrix} \xrightarrow{\div 760} \\ \xleftarrow{\times 760} \end{matrix} \text{Atmosphere} \begin{matrix} \xrightarrow{\times 101325} \\ \xleftarrow{\div 101325} \end{matrix} \text{Pascal (Pa)} \begin{matrix} \xrightarrow{\div 1000} \\ \xleftarrow{\times 1000} \end{matrix} \text{kPa} \begin{matrix} \xrightarrow{\div 100} \\ \xleftarrow{\times 100} \end{matrix} \text{Bar} \begin{matrix} \xrightarrow{\times 13.817} \\ \xleftarrow{\div 13.817} \end{matrix} \text{PSI}$$

### Gas Laws:

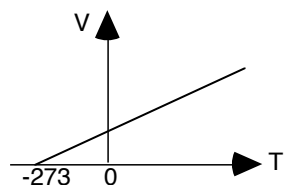
| Gas Laws: Assuming Ideal conditions  | Applications   |
|--|--|
| <p><b>Boyle's Law</b> states the Volume of gas is inversely proportional to the Pressure for a constant amount of gas at a constant temperature.</p> <p>eg. <math>V \propto \frac{1}{P}</math> so if the volume is doubled, the pressure will _____.</p>  | <p>Since <math>V \propto \frac{1}{P}</math> then <math>p \times V</math> is a constant so <math>p_1 \times V_1 = p_2 \times V_2</math>.</p> <p>eg. 500 mL of a gas at 2.0 atm pressure is <b>transferred</b> to a 2.0 L container. Determine the <b>pressure</b> of the gas in the 2.0 L container.</p> <p><math>p_1 = 2.0 \text{ atm}</math>, <math>V_1 = 500 \text{ mL} = 0.50 \text{ L}</math>, <math>V_2 = 2.0 \text{ L}</math>, Find <math>p_2</math>.</p> <p><math>p_1 \times V_1 = p_2 \times V_2</math> <span style="float: right;">The Volumes must be in same unit</span></p> <p><math>2.0 \times 0.50 = p_2 \times 2.0</math></p> <p><math>p_2 = \frac{2.0 \times 0.50}{2.0} = \mathbf{0.50 \text{ atm. Pressure.}}</math></p> <p>eg. 500 mL of a gas at 2.0 atm pressure is <b>connected</b> to a 2.0 L container at 1.0 atm pressure. Determine the resultant pressure.</p> <p><math>p_1 = 2.0 \text{ atm}</math>, <math>V_1 = 500 \text{ mL} = 0.50 \text{ L}</math>, <math>V_2 = 2.0 \text{ L}</math>, <math>p_2 = 1.0 \text{ atm}</math>, Find <math>p_3</math>.</p> <p><math>p_1 \times V_1 + p_2 \times V_2 = p_3 \times V_3</math> <span style="float: right;"><math>V_3</math> is the total volume (<math>V_1 + V_2</math>)</span></p> <p><math>2.0 \times 0.50 + 1.0 \times 2.0 = (0.50 + 2.0) \times p_3</math></p> <p><math>1.0 + 2.0 = 3.0 = 2.5 \times p_3</math></p> <p><math>p_3 = \frac{3.0}{2.5} = \mathbf{1.2 \text{ atm. Pressure.}}</math></p> |

**Charles' Law** states the Volume of gas is directly proportional to its Temperature for a constant amount of gas at a constant pressure.

$V \propto T$  so as the temp. is doubled, the volume is doubled.

However,

since calculations may involve using  $0^\circ\text{C}$  as a denominator which would lead to an undefined value, the Kelvin temperature scale is used = temperature in  $^\circ\text{C} + 273.15\text{ K}$ .



Absolute Zero =  $-273^\circ\text{C}$  is the theoretical lowest temperature attainable.

Since  $V \propto T$ , then  $\frac{V}{T}$  is a constant so  $\frac{V_1}{T_1} = \frac{V_2}{T_2}$ .

eg. 500 mL of a gas at  $25^\circ\text{C}$  is heated to  $120^\circ\text{C}$ . Determine the new **volume** the gas will occupy.

$V_1 = 500\text{ mL}$ ,  $T_1 = 25^\circ\text{C} + 273 = 298\text{ K}$ ,  $T_2 = 120^\circ\text{C} + 273 = 393\text{ K}$ . Find  $V_2$ .

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \quad \text{so} \quad \frac{500}{298} = \frac{V_2}{393}$$

$$V_2 = \frac{500 \times 393}{298} = \mathbf{659\text{ mL}}.$$

eg. 500 mL of a gas at  $20^\circ\text{C}$  is **connected** to a 2.0 L flask container at  $50^\circ\text{C}$ . Determine the resultant temperature.

$T_1 = 20^\circ\text{C} = 293\text{ K}$ ,  $V_1 = 500\text{ mL}$ ,  $V_2 = 2.0\text{ L} = 2000\text{ mL}$ ,  $T_2 = 50^\circ\text{C} = 323\text{ K}$ , Find  $T_3$ .

$$\frac{V_1}{T_1} + \frac{V_2}{T_2} = \frac{\text{Final Volume}}{\text{Final Temperature}} \quad \& \quad V_3 \text{ is the total volume } (V_1 + V_2)$$

$$\frac{500}{293} + \frac{2000}{323} = \frac{2500}{\text{Final Temperature}} \quad \text{so} \quad 7.8984 = \frac{2500}{\text{Final Temperature}}$$

$$T_3 = \frac{2500}{7.8984} = 316.5\text{ K} - \text{ } = \mathbf{43.5^\circ\text{C}}.$$

### Combining Charles' & Boyle's Laws

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2} \quad \text{for transferring gases}$$

$$\frac{P_1 V_1}{T_1} + \frac{P_2 V_2}{T_2} = \frac{P_3 V_3}{T_3} \quad \text{for connecting two flasks together}$$

"B" for Boyle comes before "C" for Charles like "P" for pressure comes before "T" for temp.

The above laws can be used for any calculations involving more than 1 set of Volumes or Pressures or Temperatures when the General Gas Equation would not be suitable.

eg. The gas in a 2.00 L flask at 101.3 kPa at  $20^\circ\text{C}$  is **transferred** to an empty 5.00 L flask at  $0^\circ\text{C}$ . Find the **pressure** in the 2nd flask.

$V_1 = 2.00\text{ L}$ ,  $P_1 = 101.3\text{ kPa}$ ,  $T_1 = 20 + 273 = 293\text{ K}$ ,  $V_2 = 5.00 + 2.00 = 7.00\text{ L}$ ,  $T_2 = 0 + 273 =$

$273\text{ K}$ . Find  $P_2$ .

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}, \quad \text{so} \quad P_2 = \frac{P_1 V_1 T_2}{T_1 V_2} = \frac{101.3 \times 2.00 \times 273}{293 \times 7.0} \quad \text{so} \quad P_2 = \mathbf{27.0\text{ kPa}}.$$

eg. A 2.00 L flask at 101.3 kPa at  $20^\circ\text{C}$  is **connected** to 3.00 L flask at 202.6 kPa &  $25^\circ\text{C}$ . Find the resultant **pressure** if the flasks are chilled to  $0^\circ\text{C}$

$V_1 = 2.00\text{ L}$ ,  $P_1 = 101.3\text{ kPa}$ ,  $T_1 = 20 + 273 = 293\text{ K}$ ,  $V_2 = 3.00\text{ L}$ ,  $P_2 = 202.6\text{ kPa}$ ,  $T_2 = 25 + 273 =$

$298\text{ K}$ ,  $V_3 = 2 + 3 = 5.00\text{ L}$ . Find  $P_3$

$$\frac{P_1 V_1}{T_1} + \frac{P_2 V_2}{T_2} = \frac{P_3 V_3}{T_3} \quad \text{so} \quad \frac{101.3 \times 2.00}{293} + \frac{202.6 \times 3.00}{298} = \frac{P_3 \times 5.00}{273}$$

$$0.69 + 2.04 = \frac{P_3 \times 5.00}{273} \quad \text{so} \quad 2.73 = \frac{P_3 \times 5.00}{273}$$

$$P_3 = \frac{2.73 \times 273}{5.00} = \mathbf{149.1\text{ kPa}}$$

**General Gas Equation** combines the laws ( $pV, \frac{V}{T}$  &  $\frac{V}{n} = \text{constant}$ ) into one equation,

**$p \times V = n \times R \times T$**  where

Pressure, p is in **kPa**,

Volume, V is in **L**,

Amount, n is in **mol**,

R is a Gas Constant,  **$8.314\text{ J K}^{-1}\text{ mol}^{-1}$**

Temperature, T is in **K** ( $^\circ\text{C} + 273\text{ K}$ ).

### Converting units:

Volume:  $1\text{ m}^3 = 1000\text{ L} = 1000\text{ dm}^3$ .

$1\text{ L (dm}^3) = \text{ } \text{mL (cm}^3)$ .

Pressure:  $1000\text{ Pa} = 1.0\text{ kPa}$ .

eg. Find the **amount** (in mol) of a gas at  $25^\circ\text{C}$  in a 250 mL flask at 120 kPa.

$V = 0.25\text{ L (250} \div 1000)$ ,  $p = 120\text{ kPa}$ ,  $T = 25 + 273 = 298\text{ K}$ . Find n.

$$n = \frac{P \times V}{R \times T} = \frac{120 \times 0.25}{8.31 \times 298}$$

$$= \mathbf{0.012\text{ mol} = 1.2 \times 10^{-2}\text{ mol}}.$$

eg. Find the **temperature** in a 2.00 L flask containing 0.10 mol of gas at 101.3 kPa..

$V = 2.00\text{ L}$ ,  $n = 0.10\text{ mol}$ ,  $p = 101.3\text{ kPa}$ . Find T.

$$T = \frac{P \times V}{R \times n}$$

$$T = \frac{101.3 \times 2.00}{8.31 \times 0.10} = 244\text{ K}$$

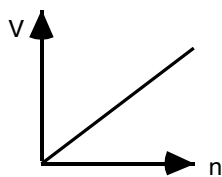
$$T = 244\text{ K} - 273\text{ K} = \mathbf{-29^\circ\text{C}}$$

Other examples can be used to determine the Pressure or Volume.

**Relating amounts to Volume:**

The volume occupied by a gas is directly proportional to the amount of gas present at a constant temperature and pressure.  $V \propto n$  so if the

amount is doubled, the volume will \_\_\_\_\_.



Since  $V \propto n$ , then  $\frac{V}{n}$  is a constant so  $\frac{V_1}{n_1} = \frac{V_2}{n_2}$ .

eg. In the reaction  $2 \text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2 \text{SO}_3(\text{g})$  there is 20.0 L of  $\text{SO}_2$  present. Determine the **volume of  $\text{SO}_3$**  produced.

$V_1(\text{SO}_2) = 20.0 \text{ L}$ ,  $n_1(\text{SO}_2) = 2$  (see reaction),  $n_2(\text{SO}_3) = 2$  (see reaction). Find  $V_2(\text{SO}_3)$ .

The reaction coefficients can substitute for the amount in mol of each gas.

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$

$$V_2(\text{SO}_3) = \frac{n_2 V_1}{n_1} = \frac{2 \times 20.0}{2} = \mathbf{20.0 \text{ L}}$$

This means that the volumes will be in the same ratios as the mol coefficients in the reaction equations. eg.  $2 \text{SO}_2 \rightarrow 2 \text{SO}_3$

20 L      20L

Check out the website; [http://www.chem.ufl.edu/~itl/2045/lectures/lec\\_d.html](http://www.chem.ufl.edu/~itl/2045/lectures/lec_d.html). Click on the **Simulation** link in the **Molecular Picture Of Pressure** section. Select the Temp. & Pressure options and slide the amounts to see the effect.

In reality, these laws are not totally accurate since at very low temperatures, the particles can still vibrate and there are still interactions between the particles.

The volume occupied by 1 mol of **any** gas at any specific temperature and pressure is called the **Molar Volume** (symbol,  $V_m$ ). It is determined using the formula,  $V_m = \frac{V}{n}$  or  $\frac{R \times T}{P}$  and has the unit L/mol.

At **Standard Temperature & Pressure (S.T.P.)**, 0 °C (273 K) & 101.325 kPa the  $V_m$  for any gas = **22.4 L/mol**.

At **Standard Laboratory Conditions (S.L.C.)**, 25 °C (298 K) & 101.325 kPa the  $V_m$  = \_\_\_\_\_ L/mol.

eg. Find the **amount** in mol of gas in a 2.2 L flask at S.L.C.

$$n(\text{gas}) = \frac{V}{V_m} = \frac{2.2}{24.45} = \mathbf{9.0 \times 10^{-2} \text{ mol}}$$

However some values of  $V_m$  are slightly different which cannot be explained by the Kinetic Theory of Gases and maybe due to the interactions between different gas molecules.

The Gas Laws can be used in Stoichiometry calculations:

eg). Find the **mass** of Ammonia,  $\text{NH}_3$  that will undergo combustion in the presence of air ( $\text{O}_2(\text{g})$ ) to produce 4.89 L of Nitrous Oxide,  $\text{NO}$  at 25 °C and 101.3 kPa and steam ( $\text{H}_2\text{O}(\text{g})$ ).

$V(\text{NO}) = 4.89 \text{ L}$ ,  $T = 25 + 273 = 298 \text{ K}$ ,  $p = 101.3 \text{ kPa}$ . Find  $n(\text{NO})$  initially.

**Step 1: Write a balanced equation (with states).**

eg. \_\_\_\_\_  $\text{NH}_3(\text{g}) +$  \_\_\_\_\_  $\text{O}_2(\text{g}) \rightarrow$  \_\_\_\_\_  $\text{NO}(\text{g}) +$  \_\_\_\_\_  $\text{H}_2\text{O}(\text{g})$ .

**Step 2: Find the amount, n(known) using measurements known about a chemical and a mol formula:**

eg.  $n(\text{NO}) = \frac{P \times V}{R \times T} =$  \_\_\_\_\_ Remember,  $P$  is in kPa,  $V$  is in L and  $T$  in K.  
= \_\_\_\_\_

**Step 3: Use the equation coefficients and n(known) to find the amount, n(want).**

$\frac{n(\text{Wanted chem})}{n(\text{Known Chem})} = \frac{\text{Coefficient of Wanted Chem}}{\text{Coefficient of Known Chem}}$  so  $n(\text{Wanted Chem}) = \frac{\text{Coefficient of Chemical Wanted}}{\text{Coefficient of Chemical Known}} \times n(\text{Known Chem})$

eg.  $n(\text{NH}_3) =$  \_\_\_\_\_  
= \_\_\_\_\_

**Step 4: Convert this mol into the required unit.**

eg.  $m(\text{NH}_3) =$  \_\_\_\_\_  
= \_\_\_\_\_

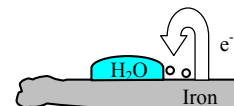
$A_r(\text{N}) = 14$ ,  $A_r(\text{H}) = 1$ .

**Congratulations on surviving Yr 11 Chemistry!**

- Pg 1:** **Water:** Photosynthesis; Oxygen Atom; Decrease; H-Bonds; Bomb Calorimeter; Releasing.  
**Solubility:** Ion-Dipole or Dipole-Dipole Interactions; Insoluble.
- Pg 2:** Hydronium; Saturated; Decreases; 100 g; 100 mL.
- Pg 3:** 1000,000 g; 184 g.  
25,000 ppb;  $0.0025 (2.5 \times 10^{-3}) \% \text{ w/w}$ ;  $0.000025 (2.5 \times 10^{-5}) \text{ g/mL}$ .  
 $n(\text{NaCl}) = 0.25 \times 0.500 = 0.13 \text{ mol}$ ,
- Pg 4:** Solvent.  
**Reactions In Solutions:** Step 1:  $\text{Pb}^{2+}$ ;  $\text{NO}_3^-$ ;  $\text{Al}^{3+}$ ;  $\text{Cl}^-$   
Step 2:  $\text{Pb}(\text{NO}_3)_2$  &  $\text{AlCl}_3$   
Step 3:  $\text{Pb}^{2+}$  &  $\text{Cl}^-$ ;  $\text{Al}^{3+}$  &  $\text{NO}_3^-$   
Step 4:  $\text{PbCl}_2$  &  $\text{Al}(\text{NO}_3)_3$  The order the products (or reactants) does not matter but the cation must be written 1<sup>st</sup>  
Step 5:  $3 \text{ Pb}(\text{NO}_3)_2(\text{aq}) + 2 \text{ AlCl}_3(\text{aq}) \rightarrow 3 \text{ PbCl}_2(\text{s}) + 2 \text{ Al}(\text{NO}_3)_3(\text{aq})$ .
- Pg 5:** Spectator; a). Calcium Carbonate, c). Silver Chloride.  
**Writing Ionic Equations:** eg.  $\text{CaCO}_3(\text{s}) + 2 \text{ H}^+(\text{aq}) \rightarrow \text{Ca}^{2+}(\text{aq}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l})$ .
- Pg 6:** Deionised  
**Introducing Acids & Bases:** Pink; Water; Water; Limestone.
- Pg 7:** Ionic Salt;  $\text{H}_2(\text{g})$ .  
**Acids & Bases:** Accept; Double Sided ( $\rightleftharpoons$ ).
- Pg 8:** Diprotic; Complete;  $1.0 \times 10^{-14}$ ; Less Than (<); 2; 0.020;  $-\log_{10}[0.020] = 1.7$ ;  $1.0 \times 10^{-3} \text{ M}$ .
- Pg 9:** Step 2:  $2.0 \times 0.200 = 0.40 \text{ mol}$ .  
Step 3:  $n(\text{HCl})$ .  
Balanced Equation.
- Pg 10:**  $n(\text{Known Chemical})$ .  
**Volumetric Analysis:** Indicator; Titre Volume.
- Pg 11:** **Redox Reactions:** Increase; Reduction;  $\text{e}^-$ .  
 $\text{ON}(\text{K}) = +1$ ,  $\text{ON}(\text{Mn}) = +7$ ,  $\text{ON}(\text{O}) = -2$ ; Reductant = Fe, Oxidant =  $\text{O}_2$ .  
**Electrochemical Cells:** Ions,
- Pg 12:** Increase; Reduction.  
Step 2:  $\text{Fe}^{2+}(\text{aq}) + 2 \text{ e}^- \rightarrow \text{Fe}(\text{s})$ .
- Pg 13:** Step 3:  $\text{Zn}(\text{s}) \rightarrow \text{Zn}^{2+}(\text{aq}) + 2 \text{ e}^-$ .  
Step 4:  $\text{Fe}^{2+}(\text{aq}) + \text{Zn}(\text{s}) \rightarrow \text{Fe}(\text{s}) + \text{Zn}^{2+}(\text{aq})$ . No factors required since each Eqn has 2  $\text{e}^-$ .  
Step 5:  $-0.44 - -0.74 = +0.30 \text{ V}$ .
- Pg 14:** Oxidation; Higher; Weakest.  
**Green Technology:** Less; Highest.  
**The Atmosphere:** UV.
- Pg 15:** Respiration.  
**Nitrogen Cycle:** -3; +5.
- Pg 16:** Respiration; Greater;  
**Gases Of The Atmosphere:** Decreases; Dissolve; Combustion.
- Pg 17:** Photosynthesis; Nitrogen Fixation; Outershell or Valence Shell; Liquid Air.
- Pg 18:** Acid Rain.  
**Physical Properties of Gases:** Straight; -273.15; Mass.  
**Gas Laws:** Halve.
- Pg 19:** 273 K; 1000.
- Pg 20:** Double; 24.5.  
Step 1:  $4 \text{ NH}_3(\text{g}) + 5 \text{ O}_2(\text{g}) \rightarrow 4 \text{ NO}(\text{g}) + 6 \text{ H}_2\text{O}(\text{g})$ .  
Step 2:  $\frac{101.3 \times 4.89}{8.314 \times 298} = 0.200 \text{ mol}$ .  
Step 3:  $n(\text{NH}_3) = n(\text{NO}) = 0.200 \text{ mol}$ .  
Step 4:  $m(\text{NH}_3) = 0.200 \times 17 = 3.40 \text{ g}$ .

**Step 1:** Iron is **Oxidised** (in an **Oxidation** reaction) by Oxygen and Water to form Iron cations. **Oxidation** involves an \_\_\_\_\_ in positive charge and  $e^-$  are on the RHS of the  $\rightarrow$ , eg.  $Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$ .

Simultaneously, Oxygen and Water are **Reduced** (in a **Reduction** reaction) by Iron to form Hydroxide ions: **Reduction** involves a \_\_\_\_\_ in positive charge and  $e^-$  are on the LHS of the  $\rightarrow$ . eg.  $O_2(g) + 2 H_2O(l) + 4 e^- \rightarrow 4 OH^-(aq)$ .



To determine the overall **Redox** equation, the half cell equations are added together so the \_\_\_\_\_ cancel.

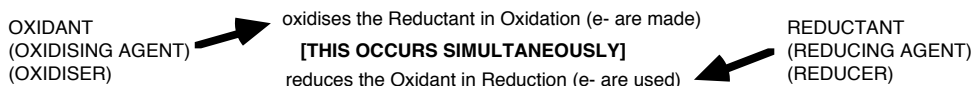
If the  $e^-$  do not cancel, each equation is multiplied by a factor so the number of electrons in both half equations are the same and cancel and the total number of atoms on both sides of the reaction arrow must also be equal.

eg. the half reaction involving Iron must be doubled. eg.  $2 Fe(s) \rightarrow 2 Fe^{2+}(aq) + 4 e^-$ .

Thus the combined reaction is:  $2 Fe(s) + O_2(g) + 2 H_2O(l) \rightarrow 4 OH^-(aq) + 2 Fe^{2+}(aq)$  that form  $2 Fe(OH)_2(s)$ .

Since Oxygen (not water) caused the Oxidation of Iron, Oxygen is called an **Oxidant**.

Since Iron caused the Reduction of Oxygen (not water), Iron is called a **Reductant**.



Water is not Reduced as there is no change in the **Oxidation Number (ON)** of either O or H which is another method used to identify Oxidation or Reduction. The Oxidation Number (O.N.) is measured for one of each atom in a molecule. For all elements, the Oxidation Number is 0 and for an ion, it is equal to the charge with the sign written 1<sup>st</sup>. eg. The ON(O in  $O_2$ ) = 0. The ON(O in  $H_2O$  or  $OH^-$ ) = -2 and the ON(H in  $H_2O$  or  $OH^-$ ) = +1. The Oxidation Number of a species undergoing Oxidation increases while it decreases in \_\_\_\_\_.

**Step 2:** The  $Fe^{2+}$  cation can be further oxidized by Oxygen and Water to form  $Fe^{3+}$ .

Oxidation:  $Fe^{2+}(aq) \rightarrow Fe^{3+}(aq) + e^-$ .

Reduction:  $O_2(g) + 2 H_2O(l) + 4 e^- \rightarrow 4 OH^-(aq)$ .

The overall Redox equation for this step is:  $4 Fe(OH)_2(s) + O_2(g) + 2 H_2O(l) \rightarrow 4 Fe(OH)_3(s)$ .

**Step 3:** Iron (III) Hydroxide changes into rust,  $Fe_2O_3 \cdot xH_2O(s)$  ( $x = 0 \rightarrow 3$ ).  $4 Fe(OH)_3(s) \rightarrow 2 Fe_2O_3 \cdot xH_2O(s)$ .

The overall equation for rusting of iron is:  $4 Fe(s) + 3 O_2(g) + 2x H_2O(l) \rightarrow 2 Fe_2O_3 \cdot xH_2O(l)$  where  $x = 0$  to 3.

There are several ways to protect the metal from corrosion:

a). **Surface Protection** which involves covering the surface of iron with a protective paint or plastic or making an alloy which produces an unreactive oxide layer.

b). **Electrochemical Protection** or **Cathodic Protection** which involves:

i). connecting Iron to the (-) cathode of a DC power supply so it's kept reduced to a metal & can't be oxidized.

ii). attaching a more reactive metal (more likely to undergo \_\_\_\_\_ than iron) to the Iron to act as a **Sacrificial Anode** which is oxidised in preference to the Iron. eg.  $Zn(s) \rightarrow Zn^{2+}(aq) + 2 e^-$ .

Lithium, Sodium, Calcium and Potassium react with water and are not used to protect other metals.

eg.  $2 Na(s) + H_2O(l) \rightarrow 2 NaOH(aq)$ . The Sodium Hydroxide (NaOH) is a strong base that can attack metals.

Other metals (eg. Magnesium and Aluminium) have a protective metal oxide layer that prevents oxidation.

### PRODUCING CHEMICALS FOR SOCIETY.

The site of chemical industries depends upon the proximity to the raw materials, transport, energy supply and workers (Chemical Engineers devise solutions to problems, monitor equipment, prepare operating budgets and determine the best means to operate the equipment using the existing conditions; Mechanical Engineers design and install equipment, solve environmental problems and help direct management; Electrical Engineers develop electrical/electronic control systems while Chemists develop analytical techniques to measure the quality of the products and together with the engineers develop ways to safely dispose of the \_\_\_\_\_ produced).

Ways of safely disposing of wastes include:

a). choosing chemicals that produce **Biodegradable Wastes** that bacteria break down in the ground. eg. paper.

b). storing **Intractable Wastes** (which cannot be made safe) in specially constructed containers that are buried deep in the ground or under the sea. eg. nuclear wastes and heavy metal compounds such as Pb, Hg and Cd.

c). incinerating **Toxic Wastes** in specially constructed ovens so only harmless products are produced which are monitored and then released into the atmosphere. eg. insecticides and other organochlorines.

### Relating Volume of Gases to their amounts:

If the question contains species in a gaseous state which are all at a specific constant temperature and pressure, then since  $pV = nRT$  and  $p$ ,  $R$  and  $T$  are now constant, the  $V \propto n$  and the mol coefficients should be in the same ratio as the volumes. This means, just use the coefficient ratios to determine the  $V$  and ignore the  $p$ ,  $R$  and  $T$  values.

eg). According to the reaction equation,  $2 \text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2 \text{SO}_3(\text{g})$ , what **volume** of Sulfur Dioxide,  $\text{SO}_2$  reacts in Oxygen to produce 8.00 L of Sulfur Trioxide,  $\text{SO}_3$  at 25 °C and 101.3 kPa.

Since the ratio of the coefficients of  $\text{SO}_2:\text{SO}_3$  is 2:2 or 1:1, the volume ratio is the same.

So if there is 8.00 L of  $\text{SO}_3$  there is **8.00 L of  $\text{SO}_2$** .

Alternatively (and taking much longer to calculate), you could use  $n(\text{SO}_3) = \frac{p \times V}{R \times T}$  to calculate the  $n(\text{SO}_3)$  and then use the equation coefficients to note the  $n(\text{SO}_2) = \frac{2}{2} \times n(\text{SO}_3)$  and then rearrange the formula,  $pV = nRT$  to find the volume of  $\text{SO}_2$  using,  $V(\text{SO}_2) = \frac{n \times R \times T}{p} = \mathbf{8.00 \text{ L of } \text{SO}_2}$ .